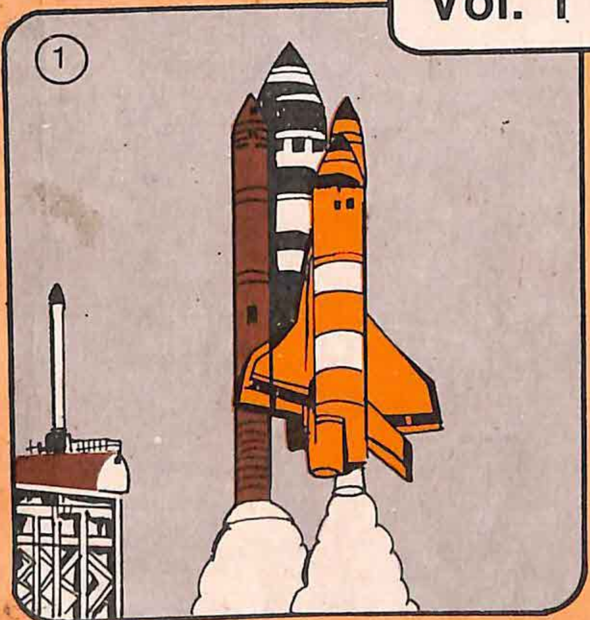
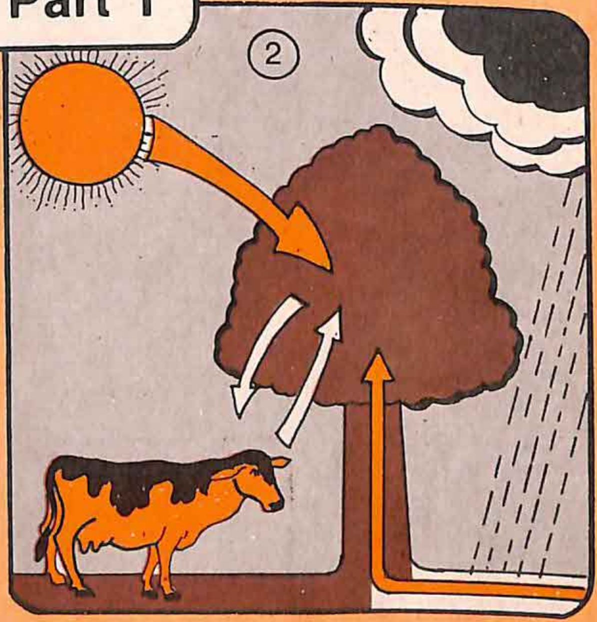
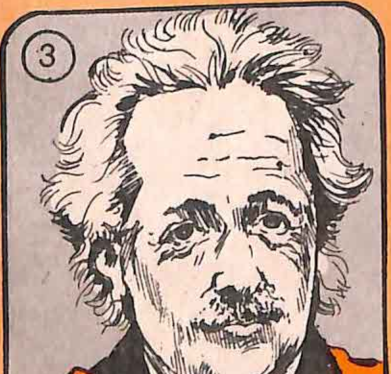
FUNDAMENTALS OF SCIENCE

Vol. I Part I







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Strictly according to the latest syllabus prescribed by the Central Board of Secondary Education, New Delhi in Science for Class IX of Secondary School Certificate Examination under the Delhi and All India Schemes.

FUNDAMENTALS OF SCIENCE

Vol. I Part I

(Covering Unit 1 : Matter—Nature and Behaviour)

(For Class IX)

By

SATISH SURI

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PREFACE

This volume covers Unit 1 of the Science syllabus for Class IX prescribed by the Central Board of Secondary Education, New Delhi for the Secondary School Examination under the Delhi and All India Schemes.

The contents of Unit 1 namely, Matter : Nature and Behaviour, are divided into six chapters to make the study of the subject systematic and easy to follow. The book contains the latest information on each topic. Each concept has been explained lucidly and illustrated with diagrams wherever possible. The language is easy to understand.

Learning objectives have been given in the beginning of each chapter. *Summary* of each chapter has been given at the end of the text for quick revision.

At the end of each chapter questions are also given which are varied in nature like *definition forms*, *objective type questions*, *fill in the blanks*, *short answer* and *essay type questions*. All the questions bear a direct relevance to the subject matter of the chapter.

I hope the book will prove useful for the students.

Suggestions for the improvement of the book will be gratefully received and acknowledged.

SATISH SURI

April, 1991

* * *

Latest Syllabus in Science prescribed for Class IX by CBSE

Unit 1 : Matter—Nature and Behaviour

Nature and behaviour of different types of substances, elements, compounds and their mixtures, Structure of matter, Atomic theory, molecules and atoms; Structure of atom—electrons, protons and neutrons; Composition of nucleus—atomic number and mass number, distribution of electrons in different energy levels in an atom, valence electrons and valency.

Atomic mass and molecular mass; mole concepts; percentage composition of compounds.

Chemical bond-formation of ionic and convalent bonds, important properties of electrovalent and covalent compounds.

Physical and chemical changes; difference between physical and chemical change; Combination reactions, displacement reactions, decomposition reactions; slow and fast reactions; catalyst; representation of chemical reactions-chemical equations; exothermic and endothermic reactions.

Electrochemical cell—construction of a simple Voltaic cell; working of an electrochemical cell; lead storage battery and dry cells.

Electrolysis—movements of ions during electrolysis; relationship between current, time and amount of metal deposited during electrolysis; electroplating.

Classification of elements—similarities and dissimilarities in the properties of some elements, Periodic law, periods and groups, gradation of properties of elements along a period and in the groups; prediction of properties of elements in the periodic table. □

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CHAPTER 1

UNIT 1 : MATTER — NATURE AND BEHAVIOUR

THE NATURE OF MATTER

LEARNING OBJECTIVES

1. Matter – its definition and examples
2. Physical and Chemical properties of a substance
3. Identifying a substance by its physical properties
4. Types of substances – Pure and Impure substances
5. Identifying the purity of a substance by its physical properties
6. Types of pure substances : Element and Compound
7. Element – Definition and its symbolic representation
8. Compound : Characteristics of a Compound
9. Formula and its importance
10. Law of constant composition and law of multiple proportion
11. Mixture – its definition and types.
12. Techniques, of industrial importance, for separating a mixture namely : Filtration ; Distillation ; Fractional distillation : Sublimation and Chromatography

INTRODUCTION

Chemistry is the science that studies the properties and reactions of substances and the laws that govern these reactions. Firstly what is meant by a substance in this context ? Substances are different kinds of matter of which all bodies and things are made of. In the present unit you will study about it in detail.

1.1 MATTER :

The stuff of which all things are made is called matter. Matter is anything that occupies volume and has mass. Mass is the amount of matter that is contained in any object. Mass and volume of an object are related to each other by a property known *Density*.

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

Wood, water, air, aluminium, paper, ash, oil, sugar and glass are a few examples of matter. Different kinds of matter are made up of different substances. The nature of one substance is different from the other. The characteristics that enable us to distinguish one substance from another are known as *properties*. Every substance possesses certain definite physical properties that help us in identifying it. These are the properties that affect sight, hearing, taste, smell and touch. Colours and smells are readily observed – Tastes are less useful because an unknown substance may be poisonous or may cause burns. Physical properties do not involve a change in the composition of the matter. In the determination of such properties either the state of the substance does not change at all or only its physical state changes. For example water from solid form changes to liquid on melting. A few physical properties are colour, hardness, ductility, malleability, physical state, melting point, boiling point, density, conductivity (electrical and thermal) etc.

Two substances may be distinguished from each other by a study of their properties. For example : Sugar and salt can be distinguished by taste ; iron and copper by colour ; and silver and mercury by physical state. In addition to physical properties a substance also exhibit chemical properties. These are the properties that involve the change in composition of the substance. For example : A chemical property is exhibited when a candle burns because the constituents of a candle combine with oxygen of air to form different substances. Rusting of iron is a chemical property because when iron rusts, iron combines with air to form a reddish brown *rust* (hydrated ferric oxide). *Although taste and smell are considered to be physical properties but they are chemical properties.*

1.2 CHEMICAL CLASSIFICATION OF MATTER :

Materials that are found in nature are either single substances or consists of two or more substances. The sample containing only one type of substance is known as *pure substance*. A pure substance is any variety of matter all specimens of which have identical chemical properties and composition. Common examples of some pure substances used in everyday life are : distilled water for car batteries ; copper used in electrical wiring ; washing soda crystals sugar. All samples of any one of these are identical. *A pure substance may be identified by its physical properties, chemical properties* and *composition* because no two substances are alike in all respects. The physical properties that can be measured exactly and are most useful are *melting point* or *freezing point* and *boiling point*. For example : A colourless, odourless liquid may be water if it freezes at 0°C and boils at 100°C and has a density of 1 gm/ml at 4°C. The physical properties can be used to test the purity of a substance. An impure substance freezes at a temperature lower than the pure substance and boils at a temperature higher than the pure substance. Sea water, common salt solution freeze below 0°C and boil above 100°C.

Substances may also be distinguished by measuring their densities because it is rare that any two substances have identical densities. One millilitre of mercury at 298 K has a mass of 13·5339 gm. So we say that density of mercury is 13·5339 gm/ml. At 273 K and 1 atmospheric pressure one litre of hydrogen has a mass of 0·08987 gm. so its density is 0·08987gm/litre.

Example 1.1. Calculate the density of a bod that has a mass of 320 gm and volume of 45 ml.

Solution :

Mass = 320 gm.

Volume = 45 ml.

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

$$= \frac{320 \text{ gm}}{45 \text{ ml}}$$

$$= 7{\cdot}1 \text{ gm/ml } \textbf{Ans.}$$

Example : 1.2. Calculate the volume occupied by 4 gm of air if the density of the air is 1·29 gm/litre.

Solution :

Mass of air = 4 gm

Density = 1·29 gm/litre

$$\text{Volume} = \frac{\text{Mass}}{\text{Density}}$$

$$= \frac{4}{1{\cdot}2 \text{ litre}}$$

$$= 3{\cdot}10 \text{ litre } \textbf{Ans.}$$

Example 1.3. The density of sulphuric acid solution that is 38% by weight of sulphuric acid is 1·30 gm/ml. Calculate the weight in grams of pure sulphuric acid present in 400 ml of this solution.

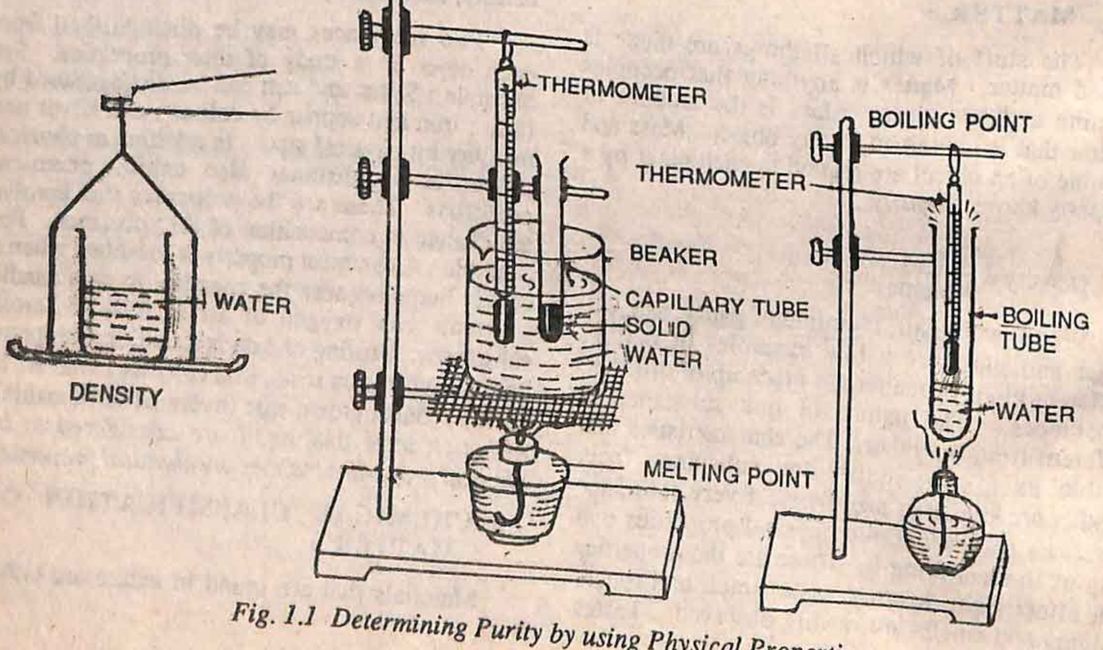

Fig. 1.1 Determining Purity by using Physical Properties

Solution : Density is 1·30 gm/ml. So 1 ml. of sulphuric acid weighs 1·30 gms. Hence 400 ml of sulphuric acid will weigh = 1·30 × 400 gms.

= 520 gms.

Since 38% by weight of acid is pure sulphuric acid. So No. of grams of sulphuric acid in 400 ml.

of solution of acid is $= \frac{520 \times 38}{100}$

= 197·6 gms. **Ans.**

1.3 ELEMENTS AND COMPOUNDS :

Pure substances are of two types :

(*a*) A chemical *Element*

(*b*) A chemical *Compound.*

Element : Elements are the basic units of all matter. These can neither be formed from simpler substances nor broken down into simpler forms of matter by chemical means. Some elements are very common. *For example* – Iron, Oxygen, Copper, Gold. This definition of an element has a limitation because the radioactive elements like uranium can be broken down into simpler elements like Barium and Krypton. So a new definition was given for elements. Elements are the substances made up of only one kind of atoms. But this definition too has a limitation due to the presence of isotopes (Isotopes are the atoms of the same element having same number of protons but different number of neutrons). Hence elements may be defined as "*Those substances that are made up of atoms having the same number of protons in their nuclei.*" Elements may be solids, liquids or gases. A few ore gases such as hydrogen, oxygen, nitrogen etc. Bromine and mercury are liquid at room temperature. Majority of the elements are solids.

It is interesting to note that the variety of substances that exist in our world today are made up of 90 elements only. So far nearly 106 elements are known. Only 90 elements occur in nature. Rest are man-made. 81 of these elements are stable. The heaviest element which is found in worthwhile amount is uranium. Some elements such as gold, platinum, copper, nitrogen etc., occur free in nature but majority of these combine with other elements to form chemical compounds or mixtures.

Elements found in the human body are : Oxygen, carbon, hydrogen, nitrogen, sulphur and phosphorus. A living cell is composed of these elements and these make up more than 99% of its weight. In addition to these, other elements present in the human body are chlorine, sodium, potassium.

Elements in Living Organisms : Many elements are accumulated in the living beings. *For example* : Iodine in sea weeds ; Vanadium in sea cucumbers and potassium is found in a few plants. Iron is found in haemoglobin of blood ; Zinc in the eyes of certain animals and magnesium in chlorophyll.

Elements in Sea : *The elements that occur in major proportion in the sea are called major elements.* These are chlorine, sodium, magnesium, sulphur, calcium, potassium and carbon. Along with these a number of other elements are present in very small amounts and are known as minor elements. Examples of these are: Bromine, boron, strontium, silicon, nitrogen, lithium, aluminium and rubidium, iron, zinc, iodine, calcium and magnesium.

Significance of a Symbol : A symbol gives us the following informations :

1. It tells the name of the element.
2. It represents an atom of the element.
3. It also stands for the atomic mass of the element as compared to one-twelfth of the mass of one atom of carbon.

For example, the symbol N stands for

(*a*) the element Nitrogen

(*b*) one atom of Nitrogen

(*c*) atomic mass of N *i.e.*, 14 a.m.u.

Elements on the Earth Crust :

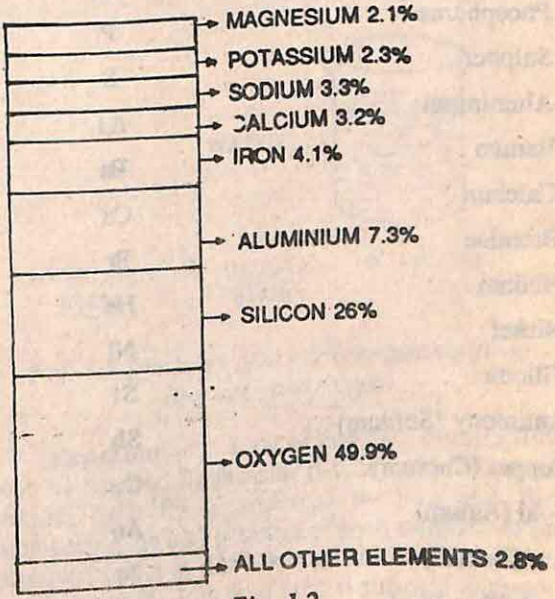

Fig. 1.2

Symbol : There is no standard system for the single substances or consists of two or more substances. The sample containing only one type of substance is known as *pure substance*. A pure substance is any variety of matter all specimens of which have identical chemical properties and composition. Common examples of some pure substances used in everyday life are : distilled water for car batteries ; copper used in electrical wiring ; washing soda crystals sugar. All samples of any one of these are identical. *A pure substance may be identified by its physical properties, chemical properties* and *composition* because no two substances are alike in all respects. The physical properties that can be measured exactly and are most useful are *melting point* or *freezing point* and *boiling point*. For example : A colourless, odourless liquid may be water if it freezes at 0°C and boils at 100°C and has a density of 1 gm/ml at 4°C. The physical properties can be used to test the purity of a substance. An impure substance freezes at a temperature lower than the pure substance and boils at a temperature higher than the pure substance. Sea water, common salt solution freeze below 0°C and boil above 100°C.

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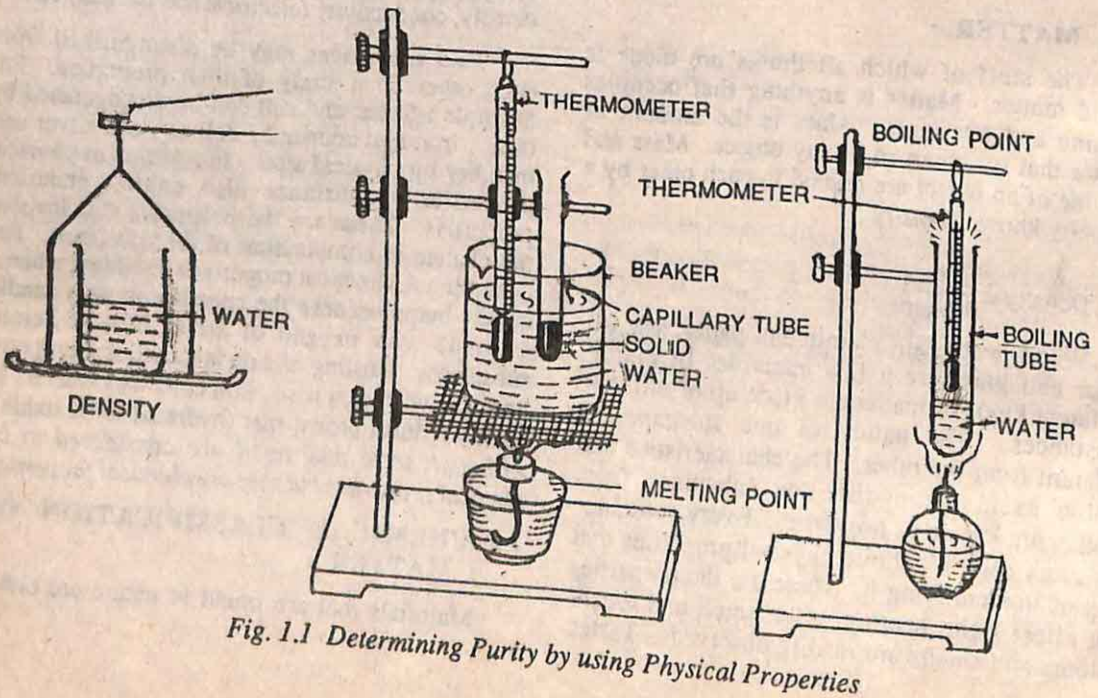

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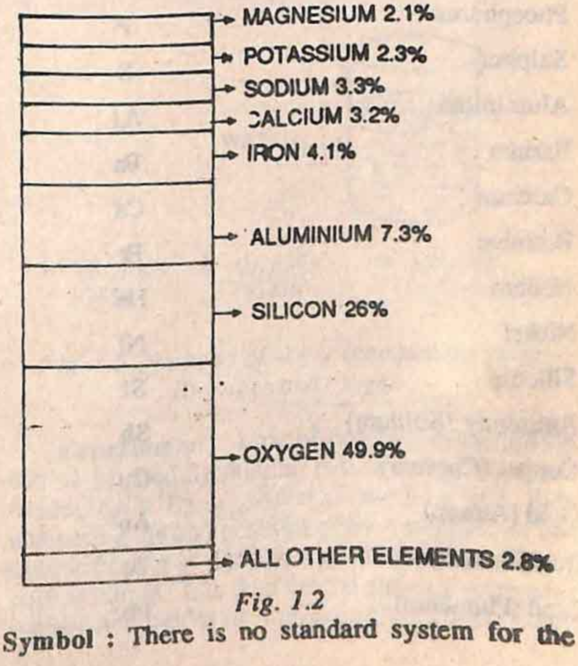


Fig. 1.2

Symbol : There is no standard system for the naming of the elements. Some were named for famous scientists ; others were named after localities or the source from which they were obtained. The name of an element varies according to the language. So a need was felt that name of each element should be represented by a *symbol* that is uniformly used throughout the world. Most of the symbols are derived from the first or two letters of the element's name. Some symbols are derived from the Latin name of the element. When a symbol contains two letters, the first is capitalised and the second is a small letter. The symbol of an element may represent among other quantities one atom of an element. Atoms may be thought of as the *structural building blocks of the elements*. Atoms are *the smallest identifiable unit of an element that can take part in a chemical change, characteristic of that element.** Symbols of some common elements are given below :

Table 1.3 Elements and their Symbols

Element	*Symbol*
Boron	B
Carbon	C
Fluorine	F
Hydrogen	H
Iodine	I
Nitrogen	N
Oxygen	O
Phosphorus	P
Sulphur	S
Aluminium	Al
Barium	Ba
Calcium	Ca
Bromine	Br
Helium	He
Nickel	Ni
Silicon	Si
Antimony (Stibium)	Sb
Copper (Cuprum)	Cu
Gold (Aurum)	Au
Iron (Ferrum)	Fe
Lead (Plumbum)	Pb
Mercury (Hydrargyrum)	Hg
Potassium (Kalium)	K
Silver (Argentum)	Ag
Sodium (Natrium)	Na
Tin (Stannum)	Sn

Note: *Names in parantheses are latin names.*

On the basis of the common properties elements can be grouped into three categories namely metal, metalloids and non-metals. 80% of the elements are classified as metals.

Metals : These elements have a lustre (*i.e.*, shine) and are good conductors. These are malleable and ductile. Examples are : Copper, gold, silver etc.

Non-metals : These are generally gases and have no lustre. These are insulators and do not exhibit malleability or ductility *e.g.*, oxygen, sulphur.

Metalloids: Elements such as silicon, germanium and arsenic have intermediate characteristics between metals and non-metals. These are characterised by slight electrical conductivity that increase with temperature.

The chemical elements can be distinguished by their *atomic mass*. Atomic mass is the quantity of matter contained in the respective atoms. For example : the lightest of the atomic masses is that of hydrogen of nearly 10^{-24} gms. However atomic weight is used as a unit, which is a number of comparing the relative weight of an atom of an element compared with that of an atom of carbon – 12 *i.e.*, the atomic weight of carbon is taken as 12 and that of any other element with respect to it. For example : Hydrogen = 1·008 ; Helium = 4·003; Oxygen = 15·999 and Uranium = 238·03 and so on.

Another important characteristic is the *atomic number* of the element. This is the number of protons present in the nucleus of an atom of that element. Atomic of number Helium = 2 : Oxygen = 8 ; Uranium = 92 and so on.

Compounds : Most of the things with which we are familiar such as fats, proteins, vitamins, sugar, water etc., are not free elements. These are formed by the combination of a few elements when two or more different elements chemically combine together in a fixed proportion, a compound is formed. The most common compound on earth is water. It is made up of two elements hydrogen and oxygen. Atoms of hydrogen and oxygen combine to give group of atoms called water *molecules.*

Molecule is the smallest particle in a compound that possesses the properties of that compound. In forming compounds, the elements lose their characteristic properties. *For example :* Sodium is a soft shining and very reactive metal that is not found free in nature. Chlorine is a greenish yellow poisonous gas and does not occur free in nature. But sodium chloride obtained by their chemical combination is used in our daily life as table salt and exist free in nature also. Hydrogen is a gas that is combustible. Oxygen, a gas, is supporter of combustion. But water formed by the chemical combination of the two is neither combustible nor supporter of combustion but is used for extinguishing the fire. Hence we can say that *the properties of a compound are quite different from the properties of the constituent elements.*

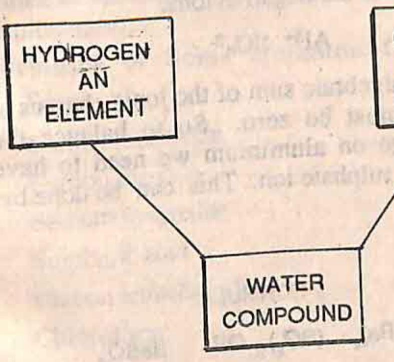


Fig. 1.3 Formation of the Compound from its Elements

1.4 LAW OF CONSTANT COMPOSITION

Dalton studied the composition of several compounds and found that all *pure samples of the same compound contains the same elements in the same proportion by mass.* This statement is known as *Law of Constant Composition. For example* : Water taken from any source is always composed of hydrogen and oxygen. The ratio of hydrogen to oxygen is always the same *i.e.*, 2 : 1 by volume.

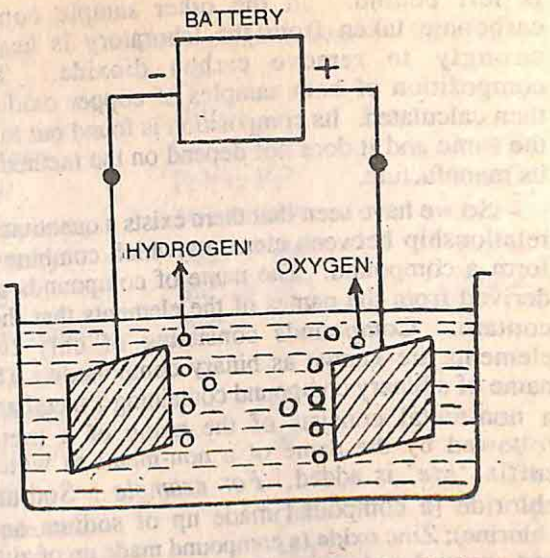


Fig. 1.4 Electrolysis of water forming hydrogen and oxygen (elements)

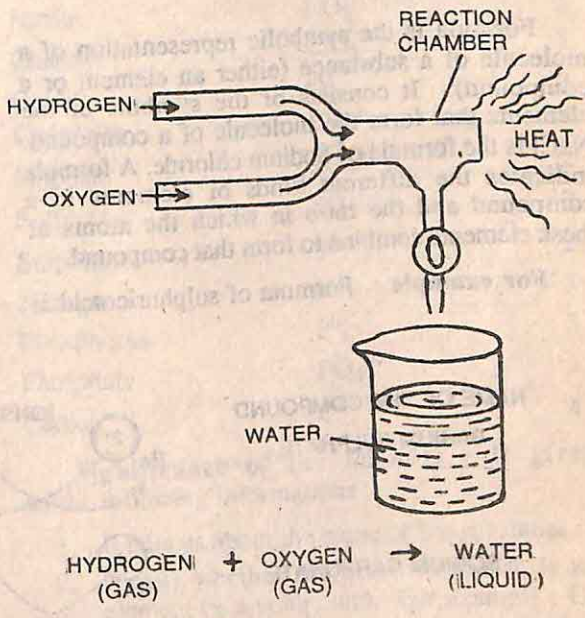


Fig. 1.5 Synthesis of water (compound) from hydrogen and oxygen

Experiment : Law of constant composition can be verified by making two samples of copper oxide, each by a different method and then comparing the composition of each sample. In one sample copper is dissolved in nitric acid to obtain copper nitrate. It is then heated strongly to remove oxygen and oxides of nitrogen. Pure copper oxide is left behind. In the other sample copper carbonate taken from the laboratory is heated strongly to remove carbon dioxide. The composition of both samples of copper oxide is then calculated. Its composition is found out to be the same and it does not depend on the method of its manufacture.

So we have seen that there exists a quantitative relationship between elements which combine to form a compound. The name of compounds are derived from the names of the elements that they contain. Compounds consisting of only two elements are known as binary compounds. The name of a binary compound containing a metal and a non-metal consists of the name of a metal followed by the name of a non-metal to which suffix '*ide*' is added. *For example :* Sodium chloride [a compound made up of sodium and chlorine); Zinc oxide (a compound made up of zinc and oxygen) and so on. The names of compounds are represented by formula.

1.5 FORMULA

Formula is the symbolic representation of a molecule of a substance (either an element or a compound). It consists of the symbols of the elements that form the molecule of a compound. NaCl is the formula of Sodium chloride. A formula indicates the different kinds of elements in a compound and the ratio in which the atoms of these elements combine to form that compound.

For example : Formula of sulphuric acid is H_2SO_4. It indicates :

1. Sulphuric acid is made up of hydrogen, sulphur and oxygen elements.
2. Atoms of these elements combine in the ratio of 2 : 1 : 4 (*i.e.*, two atoms of hydrogen, one atom of sulphur and four atoms of oxygen).

How to write the formula of a compound ?

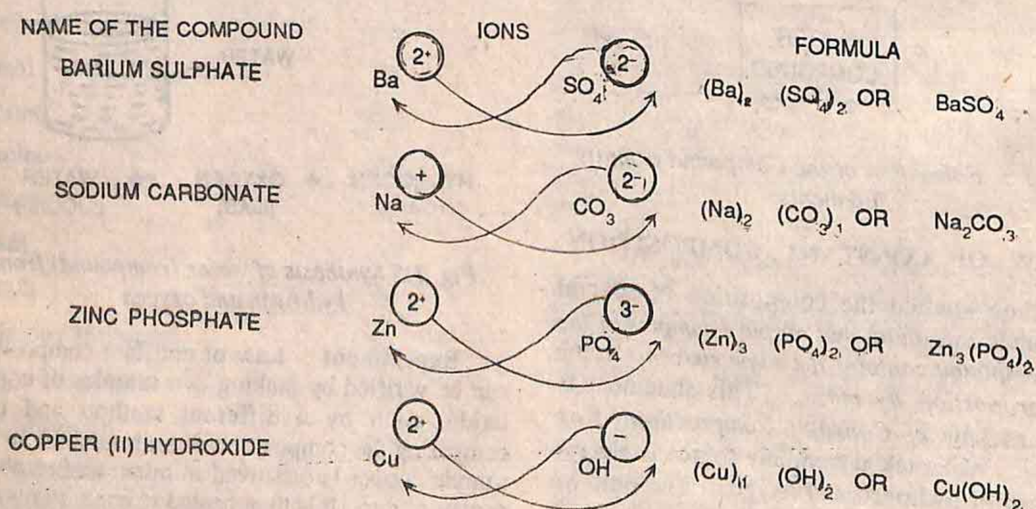
1. In order to write the formula of a compound the symbols of radicals (elements) with positive charges are written on the left-hand side and those with negative charges are written on the right-hand side.

2. The number of positive and negative charges are then equalised by putting the numbers as subscripts of the symbols of elements. This is because the algebraic sum of the ionic charges in the formula must be zero. *For example* : Formula of aluminium sulphate is written as follows :

(*a*) The symbol and charge of the positive ion is written first. It is then followed by those of the negative ions.

$$Al^{3+} \; SO_4^{2-}$$

(*b*) The algebraic sum of the ionic charges in the formula must be zero. So to balance the positive charge on aluminium we need to have more than one sulphate ion. This can be done by



taking two aluminium ions and three sulphate ions.

$$2(Al^{3+})\ 3(SO_4^{2-})$$

(*c*) Hence two aluminium ions and three sulphate ions are needed to write the formula of aluminium sulphate. The numbers are written by a subscript. The polyatomic ion is enclosed in the bracket and subscript is placed outside the bracket. So the formula of aluminium sulphate is :

$$Al_2\ (SO_4)_3$$

A few more examples are :

Calcium sulphate

Ca^{2+} SO_4^{2-} $CaSO_4$

Sodium nitrate

Na^+ NO_3^- $NaNO_3$

Zinc hydroxide

Zn^{2+} OH^- $Zn(OH)_2$

Hence we can say while writing the formula, number of charges exchange their position and are written as subscript. However it is then changed to simplest ratio.

Formulae of Some Common Compounds

Water	H_2O
Carbon monoxide	CO
Carbon dioxide	CO_2
Sodium hydroxide	$NaOH$
Sulphuric acid	H_2SO_4
Carbon tetrachloride	CCl_4
Chloroform	$CHCl_3$
Benzene	C_6H_6
Methane	CH_4
Ammonia	NH_3

Table : Valencies of a few radicals

Metals	*Corresponding Ions*	*Valences*
Potassium	K^+	1
Silver	Ag^+	1
Sodium	Na^+	1
Ammonium	NH_4^+	1
Barium	Ba^{2+}	2
Calcium	Ca^{2+}	2
Magnesium	Mg^{2+}	2
Metals	*Corresponding Ions*	*Valences*
Nickel	Ni^{2+}	2
Zinc	Zn^{2+}	2
Copper	Cu^+ or Cu^{2+}	1 or 2
Iron	Fe^{2+} or Fe^{3+}	2 or 3
Lead	Pb^{2+} or Pb^{4+}	2 or 4
Mercury	Hg_2^{2+} or Hg^{2+}	1 or 2
Tin	Sn^{2+} or Sn^{4+}	2 or 4
Aluminium	Al^{3+}	3
Non-Metals		
Chlorine	Cl^-	1
Bromine	Br^-	1
Iodine	I^-	1
Hydroxide	OH^-	1
Bicarbonate	HCO_3^-	1
Bisulphate	HSO_4^-	1
Nitrate	NO_3^-	1
Nitrite	NO_2^-	1
Chlorate	ClO_3^-	1
Oxide	O^{2-}	2
Carbonate	CO_3^{2-}	2
Sulphate	SO_4^{2-}	2
Sulphite	SO_3^{2-}	2
Sulphide	S^{2-}	2
Nitrogen	N^{3-}	3
Phosphorus	P^{3-}	3
Phosphate	PO_4^{3-}	3
Carbon	C^{4-}	4

Significance of the formula : It gives us the following informations :

1. It tells us about the name of the substance.
2. It tells whether the given substance is an element or a compound. For example : O_2 is an element – Oxygen. But H_2O^- is a compound – water.
3. It stands for one molecule of that substance.
4. It tells the various elements present in that compound.
5. It also tells the number of atoms of each type of element that combine to form that compound.
6. It also stands for the molecular mass of the substance.

For example : The formula H_2O gives us the following information :

1. It stands for water.
2. It stands for a molecule of water.
3. It tells that water consists of hydrogen and oxygen.
4. It also tells us that in water Hydrogen and Oxygen atoms combine in the ratio 2 : 1 *i.e.*, two atoms of hydrogen combine with one atom of oxygen.
5. It stands for the molecular mass of water that is $1 \times 2 + 16 = 18$ a.m.u.

1.6 LAW OF MULTIPLE PROPORTION

Hydrogen and Oxygen do not form just one compound. These can combine in two different ways to form water and hydrogen peroxide. In the same way carbon and hydrogen form a large number of compounds known as *hydrocarbons.* Consider the hydrocarbons, ethane, ethylene and acetylene. The ratio of the weights of hydrogen that combines with a fixed weight of carbon is 2 : 4 : 6 or 1 : 2 : 3. This is a simple ratio and Dalton called it '*Law of Multiple Proportions*'. This law states : '*Sometimes two elements combine to form more than one compound, when this happens, the varying weights of one element that combines with a fixed weight of the other element are in a simple ratio*'.

1.7 MIXTURE

Most materials that we come across in our daily life are mixtures of elements and compounds in varying proportions. For example, all solutions are mixtures. Ink, milk and sea water are common mixtures. Air is a mixture of oxygen, nitrogen and other gases. Brass is a mixture of copper and zinc. *A mixture is a substance that consists of two or more elements or compounds not chemically combined together.* Mixture can be classified under two heads :

(A) **Homogeneous mixture :** These are the substances in which the properties and the composition is same throughout. Homogeneous mixtures are known as *solutions*. Homogeneous mixture occur in all the th[illegible]s namely solid, liquid and gas.

Solid Solutions : A [illegible] gold used for making jewellery is a sol[illegible] of 22 parts of gold by weight in 24 parts. The rest is copper or silver. Brass is a solution of copper with zinc.

Liquid Solutions : A mixture of alcohol in water : sugar in water.

Gaseous Solution : A mixture of any two non-reacting gases namely Hydrogen and Oxygen ; nitrogen and oxygen make gaseous solution.

Heterogeneous Mixtures : The constituents of a heterogeneous mixture are not distributed uniformly. The properties and the composition of such mixtures are not uniform. In some mixtures even the constituents can be seen with naked eye *e.g.*, sand and cement, milk.

So the properties of a mixture are governed by its constituents and that the properties are variable because the relative amount of the constituents of a mixture can be varied.

Sometimes a chemist's work involves the separation of *mixtures* containing two or more substances. Methods of separation depend on the individual physical properties such as boiling point hardness, solubility of the components of a mixture since no two substances have all properties identical. Some methods for the separation of mixtures are described below :

Filtration : It is a method of separating a solid from a liquid or a solid from a gas by passing the mixtures through a filter. The filter consists of a porous material such as paper or glass wool or fine gravel. These porous materials prevent the solid particles from passing through but does not retain the liquid or gas.

Industrial applications of Filtration Process :

1. Domestic water supplies are filtered through a layer of sand and gravel in filter bed before it is finally purified.
2. In the manufacture of sodium carbonate by *Solvay process*, the precipitated sodium bicarbonate is filtered.
3. A vacuum cleaner filters dust from air. The dust particles are trapped in a fabric or paper bag.
4. In many industrial processes that require pure gases dust is removed by electrostatic precipitators. The dust particles being charged are attracted on to the wires of opposite charge.

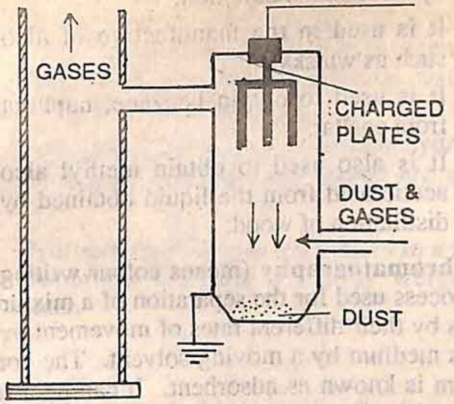


Fig 1.6 Electrostatic Precipitator, charged particles are precipitated by the electrically charged plates over which dust & gases pass.

Distillation : It is a process of boiling a liquid to form vapour and then cooling the vapours to obtain a liquid. It is used to separate a volatile substance *i.e.*, one that evaporates readily from a non-volatile substance. Impure water can be purified by distillation because ordinarily impurities in water are solids that are non-volatile.

Fractional Distillation : It is a simple distillation except that a fractionating column is attached between the distillation flask and condenser. It is used to separate two or more volatile liquids having different boiling points. A fractionating column is a tube with an irregular interior thereby providing a large cooling surface to the ascending vapours. It can be a glass column packed with leads.

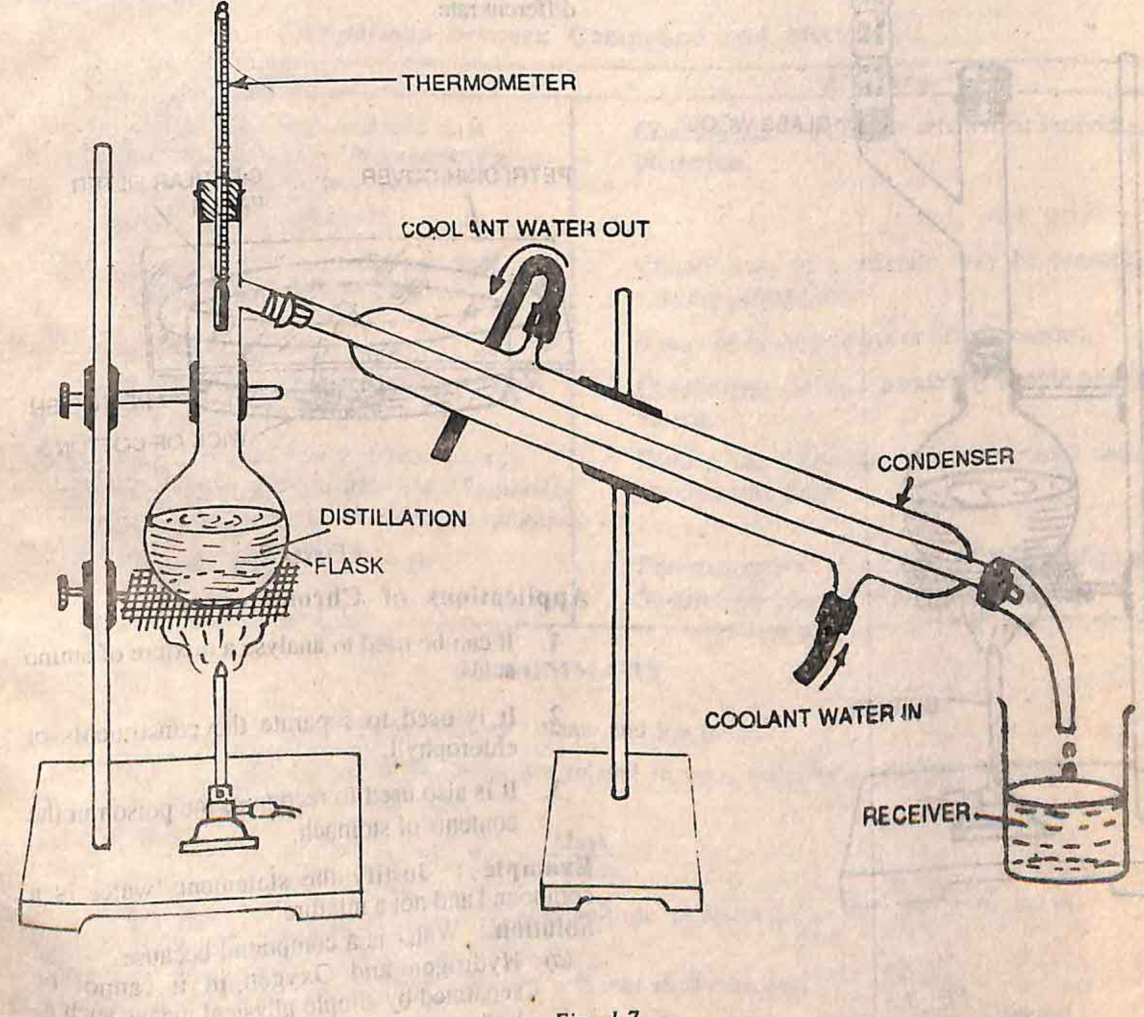


Fig. 1.7

Industrial applications of Fractional Distillation

1. Crude oil obtained from the ground is a mixture of different liquids that are separated by fractional distillation.

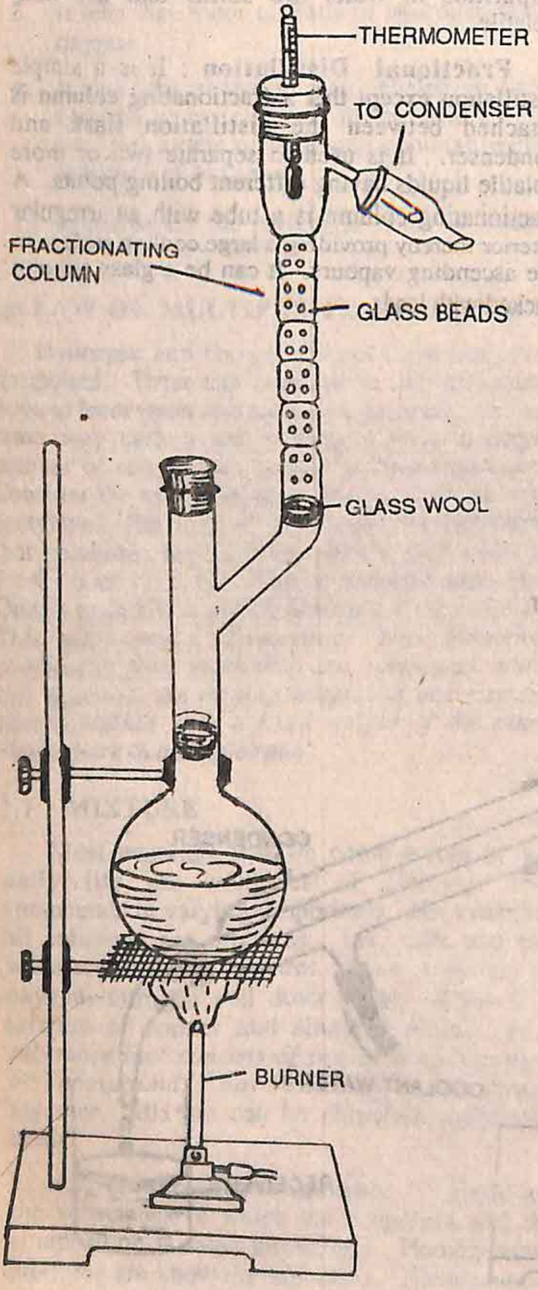


Fig. 1.8

2. Liquid nitrogen is obtained from liquid air by fractional distillation.
3. It is used in the manufacture of alcohols such as whisky.
4. It is used to obtain benzene, naphthalene from coaltar.
5. It is also used to obtain methyl alcohol, acetic acid from the liquid obtained by the distillation of wood.

Chromatography (means colour writing). It is a process used for the separation of a mixture of solutes by their different rates of movement over a porous medium by a moving solvent. The porous medium is known as adsorbent. It can be a filter paper or blotting paper. A drop of the solution of the mixture is placed near one end of the paper and it is allowed to dry. Solvent is now allowed to flow along the paper. Separation of components takes place because each pure substance moves at a different rate.

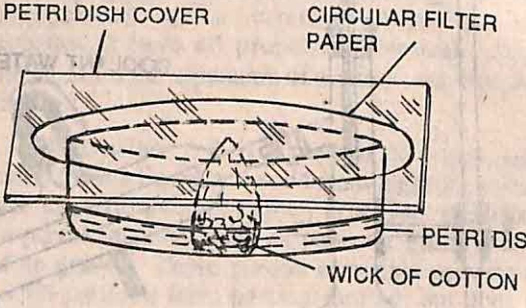


Fig. 1.9

Applications of Chromatography:

1. It can be used to analyse a mixture of amino acids.
2. It is used to separate the constituents of chlorophyll.
3. It is also used to recognise the poison in the contents of stomach.

Example : Justify the statement 'water is a compound and not a mixture'.

Solution : Water is a compound because

(*a*) Hydrogen and Oxygen in it cannot be separated by simple physical means such as boiling, distillation, filtration etc.

(*b*) Properties of water are different from its constituents *i.e.*, Hydrogen and Oxygen. Hydrogen is combustible and Oxygen is a supporter of combustion whereas water is used for extinguishing fire.

(*c*) When water is obtained by burning hydrogen in oxygen heat, light and sound are given out.

(*d*) Hydrogen and Oxygen are present in a fixed proportion of about 1 : 8 by weight in water.

Comparison between Element and Compound

Element	*Compound*
1. An element is made up of atoms having the same no. of protons in their nuclei.	A compound is made up of at least two or more different elements in fixed proportions by weight.
2. An element does not undergo decomposition in an ordinary change.	A compound undergoes decomposition in an ordinary chemical change.
3. Number of element limited (nearly 106) *For example* : Iron, Sodium.	There are an infinite number of compounds. *For example* : Water, Sodium Chloride.

Comparison between Compound and Mixture

Compound	*Mixture*
1. Constituents of a compound lose their individual properties *i.e.*, Properties of a compound are different from the elements of which it is made.	Constituents of a mixture retain their individual properties.
2. Constituents of a compound are present in a fixed proportion by weight.	Constituents of a mixture may be present in varying proportions.
3. It is homogeneous.	It may be homogeneous or heterogeneous.
4. Constituents cannot be separated by simple physical means.	Constituents can be separated by simple physical means.
5. During the formation of a mixture energy changes may or may not take place. Generally it is not accompanied by any energy change.	During the formation of a compound energy changes take place.
For example : Iron sulphide, water.	*For example* : A mixture of Iron filings and Sulphur, Hydrogen and Oxygen

SUMMARY

1.1. *Matter* is anything that occupies volume and has mass.

1.2. Mass and volume of an object are related to each other by a property known as *Density*.

$$\text{Density} = \frac{\text{Mass}}{\text{volume}}$$

1.3. Every substance possesses certain definite physical properties that help us in identifying it.

1.4. *Physical properties* do not involve a change in the composition of matter.

1.5. A pure substance is any variety of matter all specimens of which have identical chemical properties and compositions.

1.6. A pure substance has definite melting point and boiling point.

1.7. ***Elements*** are the substances that are made up of atoms having same number of protons in their nuclei.

1.8. Symbol is the shorthand representation of an atom of an element.

1.9. All the elements known today can be grouped into three categories — metals, metalloids and non-metals.

1.10. Atomic mass is the quantity of matter contained in the respective atoms.

1.11. ***Compounds*** are substances composed of two or more elements united chemically in a definite proportion by mass.

1.12. Molecule is the smallest particle of an element or a compound that is capable of independent existence and possesses the properties of that substance.

1.13. Pure samples of the same compound contains the same elements in the same proportion by mass. This is known as law of constant composition.

1.14. Formula is the symbolic representation of a molecule of an element or a compound.

1.15. When two elements combine to form two or more than two compounds then the varying weights of one element that combine with fixed weight of the other elements are in a simple ratio.

1.16. A mixture is a substance that consists of two or more elements or compounds not chemically combined together.

1.17. A homogeneous mixture is one in which the properties and composition is same throughout. Homogeneous mixture is known as solution.

1.18. Heterogeneous mixture is one in which the properties and composition is not the same throughout.

1.19. Constituents of a mixture can be separated by Filtration ; Distillation ; Fractional distillation etc.

QUESTIONS

1.1 Tick mark the correct statements :

(*i*) Pure substance is made up of only one type of substance.

(*ii*) Face cream is a pure substance.

(*iii*) Sugar solution boils at a temperature higher than pure water.

(*iv*) Sodium chloride increases the boiling point of water.

(*v*) Substances cannot be distinguished by their density.

(*vi*) Density gives the relationship between Mass and Volume of a substance.

(*vii*) The constant temperature at which, both the liquid and solid form of a substance coexist is known as freezing point of that substance.

(*viii*) Pure substance melt over a long range of temperature.

(*ix*) Glass is a pure substance.

(*x*) Pure substance is heterogeneous.

(*xi*) The most common compound on earth is water.

(*xii*) The elements retain their properties in a compound.

(*xiii*) Sodium is found free in nature.

(*xiv*) All pure samples of the same compound contains the same elements in the same proportion by mass.

(*xv*) Composition of a pure substance depends on the method of its manufacture.

(*xvi*) There is a quantitative relationship between the elements that form a compound.

(*xvii*) Formula is the symbolic representation of a molecule.

(*xviii*) The smallest particle of a substance, existing in the free state and showing specific properties is known as molecule.

(*xix*) Tap water has a fixed boiling point.

(*xx*) The algebraic sum of the ionic charges in the formulae must be zero.

1.2 Fill in the blanks :

(*i*) Elements .... be broken down into simpler species by ordinary means (can/cannot).

(*ii*) Chlorine .... an element (is/is not).

(*iii*) A substance which consists of only one type of atoms and cannot be changed into anything simpler than itself is called .... (element/pure substance).

(*iv*) An element is a pure substance made up of atoms, having the same number of .... in their nuclei (Protons/Neutrons).

(*v*) Density is equal to .... ($\frac{\text{Mass}}{\text{Volume}}$ / Mass × Volume).

(*vi*) A mixture has got a .... composition of its constitutents (fixed/variable).

(*vii*) When two or more substances are mixed physically in any ratio, the resulting product is known as .... (compound/mixture).

(*viii*) A mixture will be called .... if the composition is same throughout (homogeneous/heterogeneous).

(*ix*) A mixture .... have a definite melting point (does/does not).

(*x*) A homogeneous mixture is known as (solution/compound).

1.3 (a) Match the following :

A	B
(*a*) Glass and diamond	(*a*) Taste
(*b*) Sugar and urea	(*b*) Colour
(*c*) Graphite and copper	(*c*) Hardness
(*d*) Water and kerosene oil	(*d*) Physical state
(*e*) Silver and mercury	(*e*) Odour.

(b) Match the following :

(*a*) Lightning of bulb	(*a*) Chemical property
(*b*) Malleability	(*b*) Physical change
(*c*) Burning of coal	(*c*) Physical property
(*d*) Heat	(*d*) Matter

1.4 (*a*) Which of the following is not a physical property ?

(*a*) Colour (*b*) malleability (*c*) ductility (*d*) odour (*e*) taste (*f*) density (*g*) melting point (*h*) physical state.

(*b*) Which of the following is a pure substance?

(*a*) Rain water (*b*) Iron (*c*) Gold ornaments (*d*) Ink (*e*) Toothpaste (*f*) chalk (*g*) Ink (*h*) Milk (*i*) Sugar solution (*j*) Campa Cola.

(*c*) Which of the following elements are not found in the human body ?

(*a*) Calcium (*b*) Iodine (*c*) Gold (*d*) Nitrogen (*e*) Zinc (*f*) Sodium (*g*) Platinum (*h*) Potassium (*i*) Phosphorus (*j*) Oxygen.

1.5 (*a*) Write down the symbols of the following elements :

(*a*) Zinc (*b*) Calcium (*c*) Bromine (*d*) Potassium (*e*) Aluminium (*f*) Iron (*g*) Phosphorus (*h*) Sodium (*i*) Chlorine (*j*) Bromine

(*b*) Using the symbols and valencies of elements/radicals given in the table. Write the formulae of the following :

(*a*) Ammonium bicarbonate (*b*) Calcium carbonate (*c*) Sodium phosphate (*d*) Aluminium sulphate (*e*) Zinc nitrate (*f*) Ferrous sulphate (*g*) ferric chloride (*h*) Potassium hydroxide (*i*)Aluminium chloride (*j*) Barium hydroxide (*k*) Calcium nitrate (*l*) Barium phosphate.

1.6 (*a*) While walking on the roadside, you come across a white solid. You observed that the solid melts at 120°C and there is no change in the thermometer reading till the whole solid melts. What do you infer about the purity of this substance ? Justify your answer.

(*b*) A colourless, odourless liquid is given to you. It boils at a temperature above 100°C. Can that liquid be pure water ? Justify your answer.

Ice has a density of 0·9 gm/cm^3 whereas water has a density of 1 gm/cm^3. Will ice float on water or sink? Explain your answer.

1.7 (*a*) Name the elements that are liquidated at room temperature.

(*b*) Name the most abundant element on the earth's crust.

(*c*) Classify each of the following as element, compound or mixture :

(*a*) Wood (*b*) Carbonmonoxide (*c*) Sea water (*d*) Sugar (*e*) Blood (*f*) Air (*g*) Common salt (*h*) Iron (*i*) Steel

(*d*) Show that 'sodium chloride is a compound'.

1.8 (*a*) Define Matter. Give five examples of matter.

(*b*) Give four examples of the mixtures containing elements only.

(*c*) Write four differences between compound and mixture.

1.9 (*a*) Define compound. Support the statement 'Water is a compound' by giving two experimental evidence.

(*b*) Distinguish between :

(*a*) Element and Compound

(*b*) Homogeneous mixture and Heterogeneous mixture

(*c*) Symbol and Formula

1.10 Give statements for the followings :

(*a*) Law of constant composition (*b*) Law of multiple proportions

(*c*) Solution (*d*) Symbol.

1.11 Give two reasons to show that 'Air is a mixture and not a chemical compound'.

1.12 What are homogeneous and heterogeneous mixtures ? Give four examples of each.

1.13 How will you show that mercury is an element and mercuric oxide is a compound ?

1.14 What is fractional distillation ? What is the use of fractionating column in fractional distillation ?

1.15 Which one of the following is a homogeneous mixture ?

(*a*) A sugar solution

(*b*) Suspension of slaked lime in water.

Give reasons in support of your answer.

1.16 How will you show whether the liquid provided to you is pure or not ?

1.17 Give three reasons to show that ammonia is a compound of nitrogen and hydrogen and not a mixture of two gases.

1.18 Write short notes on :

(*a*) Filtration (*b*) Distillation (*c*) Chromatography

1.19 Give three reasons to prove that 'Air is a mixture and not a chemical compound'.

1.20 In each case describe briefly how you would carry out the following :

(*a*) Obtain pure water from sea water.

(*b*) Obtain hydrogen from a mixture of hydrogen and chlorine.

(*c*) Obtain alcohol from a mixture of alcohol and water.
(*d*) Obtain the coloured dyes in black ink.

1.21 A substance 'A' on heating gives B and C. What is A : an element, a compound or mixture ?

1.22 Two substances A and B when brought together form C with the evolution of heat. The properties of C are entirely different from those of A and B. Giving reasons explain whether C is an element or a compound or mixture.

1.23 Draw well labelled diagrams of the following :
(*a*) Distillation (*b*) Fractional distillation

PROBLEMS

1.1. Calculate the volume occupied by 2 gram of air if the density of the air is 1·29 gm/litre.
Ans : 1·55 litre

1.2. Mercury has a density of 13·5339 gm/ml. Calculate the volume occupied by 13533·9 gm. of mercury.
Ans : 1 litre

1.3. The density of sulphuric acid solution that is 38% by weight of sulphuric acid is 1·30 gm/ml. Calculate the weight in grams of pure sulphuric acid present in 500 ml. of this solution.
Ans : 247 g

1.4. Calculate the density of a body that has a 320 gms and volume of 50 ml. **Ans:** 6·2 g/ml.

1.5. At 273 K and 1 atm. pressure one litre of hydrogen has a mass of 0·08987 gm. What will be the density of hydrogen at 273 K and 1 atmospheric pressure ? **Ans ;** 0·08987 g/litre.

□□

CHAPTER 2

STRUCTURE OF ATOM

LEARNING OBJECTIVES

1. Describe, with diagrams, the cathode ray experiments.
2. List the properties of cathode rays.
3. Describe formation of positive rays.
4. List the characteristics of anode rays/positive rays.
5. Name the fundamental particles present in an atom.
6. List the properties of fundamental particles.
7. Describe Rutherford's model of atom.
8. Explain the discovery of neutrons.
9. Define atomic number and mass number.
10. State the relationship between atomic number and mass number.
11. Describe Bohr's model of atom.
12. State the rules governing the distribution of electrons in the different energy levels of an atom.
13. Write the electronic configuration of first twenty elements.
14. Define the terms valence electrons and isotopes.
15. Explain how the existence of isotopes explain the fractional atomic masses.
16. Explain the importance of X-rays.
17. Define radioactivity.
18. List the characteristics of Alpha, Beta and Gamma rays.
19. Explain the importance of radioactive isotopes.

INTRODUCTION :

So far we have seen that matter is either a pure substance or is composed of mixtures of pure substances. We have also studied the ways by which a mixture can be separated into pure substances. On the basis of results of large number of simple experiments, Dalton suggested that all matter is composed of very small particles called atoms. Further experimentation led to the belief that these atoms were solid indivisible particles. But Faraday's Laws of electrolysis suggested that certain compounds contained charged particles called ions. It was also observed that atoms of elements with high atomic number decomposed spontaneously and emitted Alpha and Beta particles. From where do these particles come from ? So it is clear that atom has a complicated structure. What is this structure ? What is atom made up of ? In this chapter we will study all about it.

2.1 DALTON'S ATOMIC THEORY :

Thousands of years ago, Kanada, the great Indian Philosopher proposed that all matter is constituted of small particles called 'Paramanu'. In 1808, John Dalton, an English school teacher, based on his experimental observations proposed *Atomic Theory* of matter. The main postulates of Dalton's Atomic Theory are as follows :

1. All matter is made up of extremely small particles called *atoms*.
2. Atoms of the same element are all alike but differ from the atoms of other elements. This is suggested by the fact that different elements have different behaviour but different samples of the same element have the same behaviour.
3. Atoms are indivisible.
4. Each atom has a definite weight.
5. Atoms of more than one element combine to form compound atoms called *molecules*.

This theory failed to explain the following electrical phenomenon :

(*a*) Different materials when rubbed together generate electrical charge.

(*b*) There are only two types of charge which scientists call positive and negative charge.

In later years the nature of the atom itself, its electrical properties and the idea that atoms might not be the ultimate particle of matter came into light.

2.2 DISCOVERY OF ELECTRONS [*Cathode Rays*] :

Under normal conditions gases are poor conductors. J.J. Thomson passed electric discharge though gases at a very very low pressure and observed a green glow on the glass. This was due to the bombardment of the glass by rays that travelled in straight line.

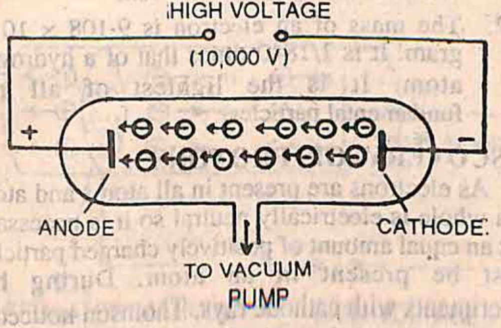


Fig 2.1 Formation of Cathode rays. Electrons are flowing from cathode to anode in a gas discharge tube.

These rays emerge from cathode and move towards anode. Since these rays originate from the cathode hence are called cathode rays. These rays were found to possess the following characteristics :

1. They cause green fluorescence on the walls of the glass tube.
2. *Cathode rays travel in a streight line* : They cause a sharp shadow, of the solid object that is placed in their path, upon the wall opposite to the cathode.

This shows that cathode rays travel in a straight line.

3. *Cathode rays produce heating effect* : These rays when made to fall on a metal foil, produce heating effect. This shows that cathode rays cause a heating effect.

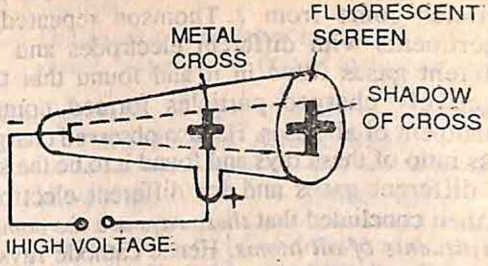


Fig 2.2 Cathode rays travel in straight lines and cast shadow of the object placed in their path upon the wall opposite the cathode.

4. *Cathode rays consist of particles having kinetic energy :* If a small paddle wheel is placed in their path, the wheel rotates. This indicates that cathode rays consist of a stream of particles having kinetic energy.

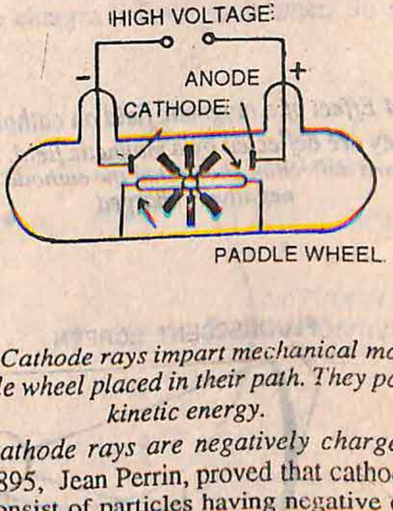


Fig. 2.3 Cathode rays impart mechanical motion to a paddle wheel placed in their path. They possess kinetic energy.

5. *Cathode rays are negatively charged :* In 1895, Jean Perrin, proved that cathode rays consist of particles having negative charge. A magnet was placed near the tube and it was observed that cathode rays were deflected in a direction corresponding to the negative charge on the particles.

Later in 1897 Thomson showed that these rays when passed through a strong electric field experience deflection. They move towards positively charged plates. This shows that the cathode rays carry negative charge. It was now clear that cathode rays are *negatively charged particles* *that travel in* straight line. To start with we had only the gas atoms, then from where did these particles come from ? Thomson repeated the experiments with different electrodes and with different gases filled in it and found that these negatively charged particles formed common constituent of all gases. He also observed charge to mass ratio of these rays and found it to be the same for different gases and for different electrodes. He then concluded that *these rays are the common constituents of all atoms*. Hence cathode rays are negatively charged particles that travel in a straight line from cathode to anode. *G.J. Stoney* called these rays as *electrons*

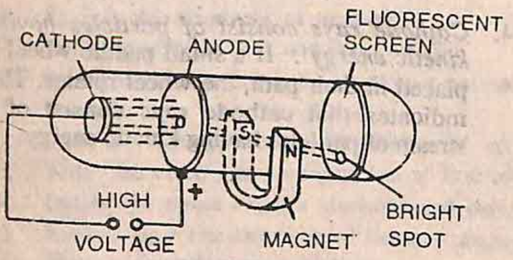


Fig. 2.4 Effect of a magnetic field on cathode rays. They are deflected by a magnetic field. The downward deflection shows that the cathode rays are negatively charged.

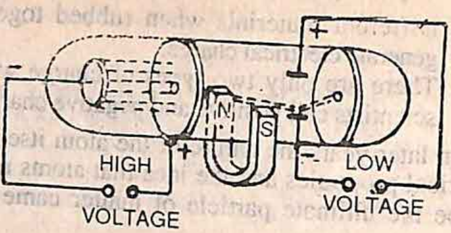


Fig. 2.5 An assembly for measurement of e/m for the particles constituting cathode rays

CHARACTERISTICS OF ELECTRONS :

1. An electrons carries a unit negative charge. The charge on an electron is $1{\cdot}6 \times 10^{-19}$ coulomb or $4{\cdot}803 \times 10^{-10}$ *esu.*
2. The mass of an electron is $9{\cdot}108 \times 10^{-28}$ gram. It is 1/1840 times that of a hydrogen atom. It is the lightest of all the fundamental particles.

DISCOVERY OF PROTONS :

As electrons are present in all atoms and atom as a whole is electrically neutral so it is necessary that an equal amount of positively charged particles must be present in an atom. During his experiments with cathode rays, Thomson noticed a red glow around the cathode, on the opposite side to the anode. In order to investigate the red glow he used a central cathode that had a hole in it as shown in the figure.

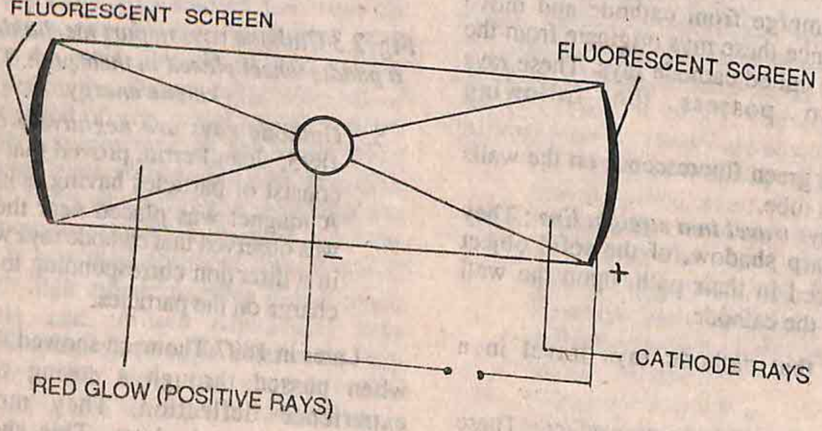


Fig. 2.6 Apparatus used by Thomson to obtain positive rays

The red glow was caused by rays which were deflected in the opposite direction to the electrons by magnetic and electric field. This showed that they contained positively charged particles. Thomson called these rays as *positive rays* *Goldstein* repeated Thomson experiment with perforated cathode. He observed the rays passing through the holes of the cathode and moving in a direction opposite to the cathode rays. On applying a strong electric or magnetic field these rays were found to deflect towards negative field. This showed that these are composed of positively charged particles. Because of positive charge on these rays they were called *positive rays* or *anode rays* (since they move towards cathode). It was found that if different gases are filled in the cathode rays tube then charge to mass ratio is not the same for all gases. It is maximum for hydrogen.

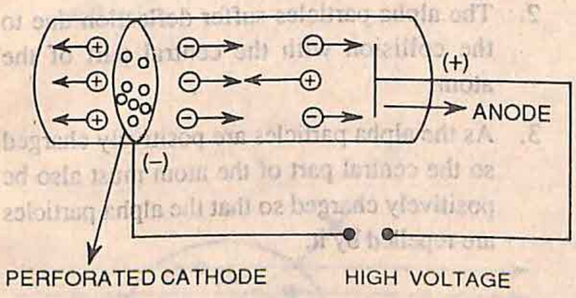


Fig. 2.7 Formation of canal rays

Characteristics of Anode Rays or Positive Rays :

1. Like cathode rays, they travel in straight line.
2. They ionise the gas through which they are passed.
3. The nature of positively charged particles depend upon the nature of the gas in the discharge tube.
4. Charge to mass ratio is different for different gases.

These positively charged particles are believed to be formed due to the knocking out of one or more electrons from a neutral atom by the fast moving electrons emitted from the cathode. These positively charged particles were called *protons*.

CHARACTERISTICS OF PROTONS

1. These carry positive charge. The positive charge on the proton is of exactly the same magnitude as the negative charge on the electron.
2. The mass of the proton is 1.673×10^{-24} grams. It is equal to that of hydrogen atom.

Comparison of the characteristics of Cathode and Anode Rays

Cathode rays	*Anode rays*
1. Composed of negatively charged particles.	Composed of positively charged particles.
2. Travel in straight line.	Travel in straight line.
3. Charge to mass ratio is same for all gases.	Charge to mass ratio different for different gases.
4. Possess kinetic energy and momentum.	Possess kinetic energy and momentum.

THOMSON'S MODEL OF ATOM

Thomson suggested that atoms consist of rings of negatively charged electrons embedded in a sphere of positive charge. The negative and positive charges balance each other. So atom is neutral.

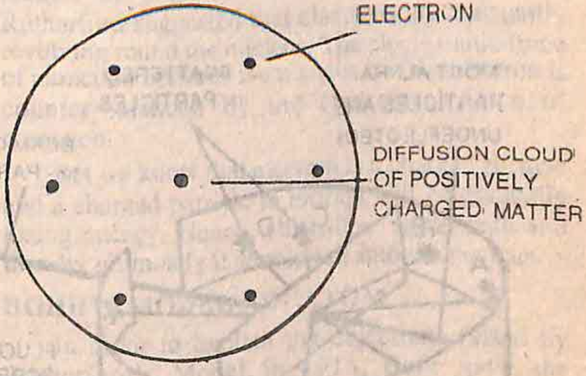


Fig. 2.8 Thomson's Model of the Atom

DISCOVERY OF NEUTRONS

So you have seen that atom is made up of protons and electrons. Hydrogen atom contains one proton and helium atom contains two protons. If there are only protons and electrons in an atom then the relative atomic mass of Helium should be twice that of hydrogen. But in fact the relative atomic mass of helium is four times the mass of hydrogen atom and not twice. It was *James Chadwick* who showed from where the extra mass in helium atoms came from. He established the existence of neutral particles, called *Neutrons*, by the bombardment of a thin sheet of beryllium with high alpha particles. Hydrogen atom has one proton, one electron and no neutron ; Helium atom has two protons, two electrons and two neutrons. Since the mass of an electron is negligible as compared to the mass of a proton and a neutron so helium atom is nearly four times as heavy as hydrogen atom.

2.3 CHARACTERISTICS OF NEUTRONS

1. It has no charge *i.e.*, is electrically neutral.
2. It has mass equal to that of a proton.

Rutherford's Experiment : (*The scattering of alpha particles*). How were the electrons arranged within the atom ? Rutherford had the idea of probing inside the atom using alpha particles from radio-active substances. He bombarded a thin strip of gold with alpha particles and determined the subsequent path of the alpha particles. A fluorescent screen was set up around the foil. Rutherford expected that most of the alpha particles might pass straight through the foil or be deviated a little. He made the following observations :

1. Majority of the alpha particles went through the foil undeflected.
2. Only a small fraction of the alpha particles were deflected by small angles and very few were deflected through large angles.
3. One in about 20000 particles *i.e.*, very very few alpha particles came straight back suffering a deflection of 180°.

So the above scattering experiment disproved the Thomson model. On the basis of Thomson's model the particles should be deflected through small angles only. This is because the positive charge in Thomson model is spread uniformly over the entire volume. Keeping the above observations in mind Rutherford reasoned as follows :

1. Since most of the alpha particles pass through the gold strip without suffering any deflection so *there must be a lot of empty space within an atom.*
2. The alpha particles suffer deflection due to the collision with the central part of the atom.
3. As the alpha particles are positively charged so the central part of the atom must also be positively charged so that the alpha particles are repelled by it.

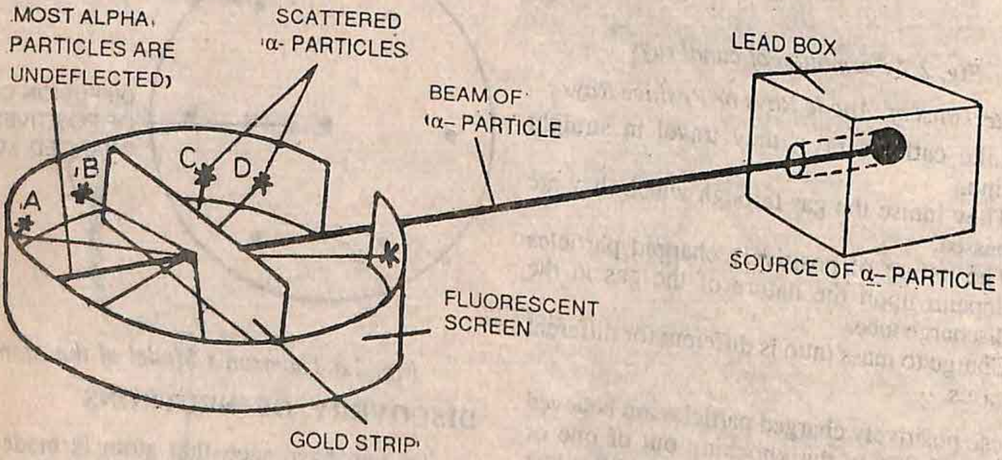


Fig. 2.9 Rutherford's α-particles scattering experiment

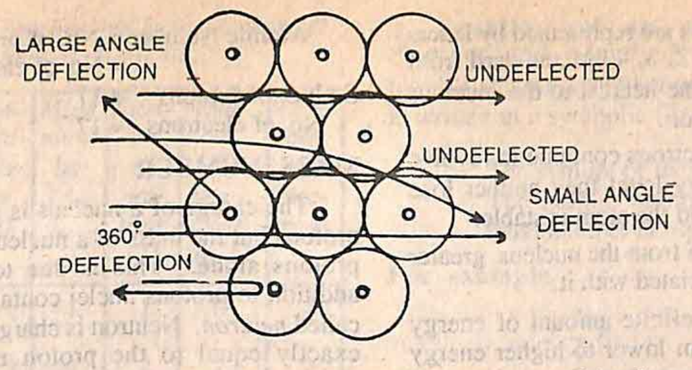


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Fig. 2.10 The general pattern of scattering of α-particles by a gold strip made up of atoms

4. As very few alpha particles suffer a large deflection and very few particles are bounced back in the direction from which they come so the whole positive charge is contained in a very small space. And whole mass of the atom is concentrated in that space. This is because kinetic energy of the alpha particles is very large and only a massive particle with high positive charge could cause this deflection.

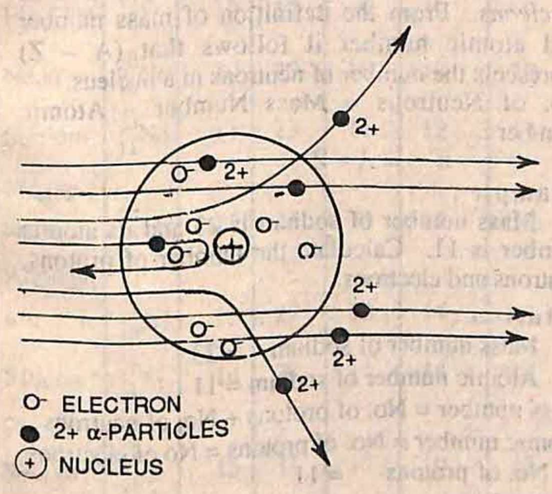

Fig 2.11 Results of the experiments that is expected on the basis of Rutherford's Model

On the basis of his experiments Rutherford put forward his model of an atom.

1. There is a positively charged centre in an atom called *Nucleus*. The mas of an atom is due to the mass of the nucleus alone.
2. The electrons are revolving round the nucleus.

The size of the nucleus is very small as compared to the size of an atom. The atomic size is about 10^5 times larger than the nuclear size.

2.4 DRAWBACK OF RUTHERFORD'S MODEL OF ATOM

According to Rutherford atom may be considered to be made up of two parts :

1. The positively charged central nucleus where the total mass of atom is concentrated.
2. The extra nuclear part containing the electrons.

So due to the forces of attraction between negatively charged electrons and the positively charged nucleus electrons should fall into the nucleus. The atom should be highly unstable. But on the other hand, we know an atom is quite stable. In order to overcome this objection Rutherford suggested that electrons are constantly revolving round the nucleus. The electrostatic force of attraction between the nucleus and the electron is counter-balanced by the centrifugal force of attraction.

But we know that electron is a charged particle and a charged particle in motion will be gradually losing energy. Hence will follow spiral path and thereby ultimately it should fall into the nucleus.

BOHR'S MODEL OF ATOM :

In order to explain the objections raised by Rutherford's Model in 1913, Bohr gave the following postulates about an atom :

1. The nucleus of an atom is situated at its centre.
2. The electrons in an atom move around the nucleus in definite circular paths called *Energy levels.*
3. Each energy level is associated with a definite amount of energy.

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4. These energy levels are represented by letters K, L, M, N, or 1, 2, 3, 4 etc.,outward from the nucleus *i.e.*, one nearest to the nucleus is K (or 1) and so on.
5. As long as the electrons continue to revolve in the same energy level they neither lose nor gain energy and the atom is stable.
6. Farther an electron from the nucleus, greater is the energy associated with it.
7. By absorbing a definite amount of energy electron jump from lower to higher energy level. It emits the same amount of energy when it comes back to that energy level.

2.5 ATOMIC NUMBER

The positive charge on the nucleus of an atom can be calculated by counting the number of alpha particles scattered in different directions. It was found that the charges of different nuclei are always integral multiple of the electron charge but with opposite sign. So if the electron charge is designated as –e, the nucleus of hydrogen atom is found to have charge +1e, the nucleus of oxygen atom + 8 e ; the nucleus of sodium atom +11e and that of uranium atom + 92e. This integer value is known as *atomic number* and is represented by the symbol Z. Since the positive charge of a nucleus is due to the positive charged particles called protons, hence *atomic number of an atom of an element is equal to the number of protons present in its nucleus.* This means atomic number of hydrogen is equal to 1 ;

Na = 1 ; Cl = 17 and so on.

Since in an atom

No. of protons = No. of electrons

So Atomic Number = No. of protons = No. of electrons.

So we can say an atom of hydrogen contain one proton in its nucleus and one electron revolves round the nucleus. An atom of sodium contains 11 protons in its nucleus and 11 electrons revolve round the nucleus. An atom of chlorine contains 17 protons in its nucleus and 17 electrons revolve round the nucleus.

Atomic Number determines the chemical identity of an atom. *Moseley* suggested that atomic number was more significant than the relative atomic mass in determining the properties of the elements.

Example :

Atomic Number of Fluorine is 9. Calculate the number of protons and electrons in its atom.

Solution : We know

Atomic Number = No. of protons
= No. of electrons
So No. of protons = 17
No. of electrons = 17

MASS NUMBER :

The charge of a nucleus is entirely due to the protons but the mass of a nucleus is not due to the protons alone. This is due to the fact that in addition to protons nuclei contain another particle called *neutron*. Neutron is chargeless and has mass exactly equal to the proton mass so the total number of protons and neutrons present in the nucleus of an atom determines the nuclear mass. Since the mass of an electron is negligible as compared to the mass of a proton or neutron so mass of an atom is equal to the nuclear mass. *This sum of protons and neutrons present in the nucleus of an atom is known as mass number.* It is represented by A.

Mass Number (A) = No. of protons + No. of neutrons

$$A = n + p.$$

Since No. of protons = Atomic Number (Z)

Hence $A = n + Z$

Protons and Neutrons are collectively known as *Nucleons.* From the definition of mass number and atomic number it follows that (A – Z) represents the number of neutrons in a nucleus.

No. of Neutrons = Mass Number – Atomic Number

$$n = A - Z$$

Example :

Mass number of sodium is 23 and its atomic number is 11. Calculate the number of protons, neutrons and electrons.

Solution :

Mass number of sodium = 23
Atomic number of sodium = 11
Mass number = No. of protons + No. of neutrons
Atomic number = No. of protons = No.of electrons
So No. of protons = 11
No. of electrons = 11
No. of neutrons = 23 –11 = 12. **Ans.**

Example :

If the atomic number of an element M is 12 and atomic mass number is 24, determine the number of neutrons in it.

Solution :

Atomc Mass Number = 24
Atomic Number = 12
No. of neutrons =Mass No. – Atomic No.
= 24 – 12 = 12 **Ans.**

Table 2.1 Atomic number and mass number of some common elements

Element	*Notation*	*Atomic number (Z)*	*Mass number (A)*	*Number of protons (Z)*	*Number of neutrons (A–Z)*	*Number of electrons*
Hydrogen	$^{1}_{1}H$	1	1	1	0	1
Helium	$^{4}_{2}He$	2	4	2	2	2
Lithium	$^{7}_{3}Li$	3	7	3	4	3
Beryllium	$^{9}_{4}Be$	4	9	4	5	4
Boron	$^{11}_{5}N$	5	11	5	6	5
Carbon	$^{12}_{6}O$	6	12	6	6	6
Nitrogen	$^{14}_{7}N$	7	14	7	7	7
Oxygen	$^{16}_{8}O$	8	16	8	8	8
Fluorine	$^{19}_{9}F$	9	19	9	10	9
Neon	$^{20}_{10}Ne$	10	20	10	10	10
Sodium	$^{23}_{11}Na$	11	23	11	12	11
Magnesium	$^{24}_{12}Mg$	12	24	12	12	12
Aluminium	$^{27}_{13}Al$	13	27	13	14	13
Silicon	$^{28}_{14}Si$	14	28	14	14	14
Phosphorus	$^{31}_{15}P$	15	31	15	16	15
Sulphur	$^{32}_{16}S$	16	32	16	16	16
Chlorine	$^{35}_{17}Cl$	17	35	17	18	17
Argon	$^{40}_{18}A$	18	40	18	22	18
Potassium	$^{39}_{19}K$	19	39	19	20	19
Calcium	$^{40}_{20}Ca$	20	40	20	20	20

Notation of an element with Mass Number and Atomic Number: The atomic number and mass number of an atom of an element is written in a symbolic form as follows :

1. The symbol of the atom is written first.
2. The atomic number is subscript, while mass number is superscript.

For example :

$$^{A}_{Z}X$$

Where X is the symbol of an element
A is Mass Number
Z is atomic number.

Magnesium – atomic number – 12 and mass number –24 is represented as

$$\boxed{^{24}_{12}Mg}$$

Example : How many protons, neutrons and electrons are there in each of the following :

(*a*) $^{4}_{2}He$

(*b*) $^{7}_{3}Li$

Solution : (*a*) $^{4}_{2}He$

Superscript indicates mass number hence mass number = 4

Subscript indicates atomic number hence atomic number = 2

Since Atomic Number = No. of protons
= No. of electrons
So No. of protons = 2
No. of electrons = 2
No. of neutrons = Mass No.–Atomic No.
= 4 – 2
= 2. **Ans.**

(*b*) $^{7}_{3}Li$

Superscript indicates of Mass No.
Hence Mass No. = 7
Subscript indicates Atomic No.
Hence Atomic No. = 3
Atomic Number = No. of protons
= No. of electrons
So No. of protons = 3
No. of electrons = 3
No. of neutrons = Mass No.–Atomic No.
= 7 – 3
= 4. **Ans.**

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2.6 DISTRIBUTION OF ELECTRONS IN DIFFERENT ENERGY LEVELS

The nucleus of an atom has not much to do with the chemical activity. Chemical activity is due to *extra nuclear electrons* orbitting the nucleus at a definite distance from the nucleus in circular paths called *Energy levels*. An energy level near to the nucleus has lowest energy and the relative energy increases as the distance from the nucleus increases. The electrons are arranged in different energy levels according to set rules which are given by Bohr – Bury Scheme as follows :

1. The maximum number of electrons that can be accommodated in each energy level is $2n^2$. n - represents the number of energy level and can have the values 1, 2, 3, 4 and so on. (Energy levels were previously known as shells or orbits). These energy levels are also represented by letters K, L, M, N etc.

So the number of electrons in the 1st energy level (K) = $2 \times 1^2 = 2$

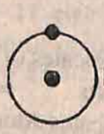
No. of electrons in 2nd energy level (L) = $2 \times 2^2 = 2 \times 4 = 8$

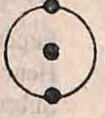
No. of electrons in 3rd energy level (M) = $2 \times 3^2 = 2 \times 9 = 18$ and so on.

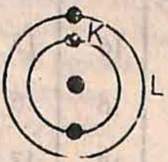
2. The outermost energy level cannot have more than 8 electrons [Except the first energy level that can have at the most 2es)
3. The penultimate energy level (*i.e.*, last but one) cannot have more than 18 electrons.
4. It is not necessary for the energy level to be completed (according to Rule No. 1) before another commences to be formed. In fact a new energy level starts filling when the outermost energy level attains 8 electrons.

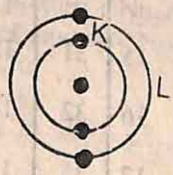
Table 2.2 Distribution of electrons for the first twenty elements

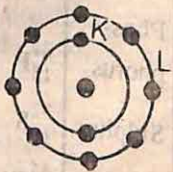
Element (symbol)	*Atomic number*	*Distribution of electrons*			
		K (n=1)	*L (n=2)*	*M (n=3)*	*N (n=4)*
H	1	1			
He	2	2			
Li	3	2	1		
Be	4	2	2		
B	5	2	3		
C	6	2	4		
N	7	2	5		
O	8	2	6		
F	9	2	7		
Ne	10	2	8		
Na	11	2	8	1	
Mg	12	2	8	2	
Al	13	2	8	3	
Si	14	2	8	4	
P	15	2	8	5	
S	16	2	8	6	
Cl	17	2	8	7	
A	18	2	8	8	
K	19	2	8	8	1
Ca	20	2	8	8	2



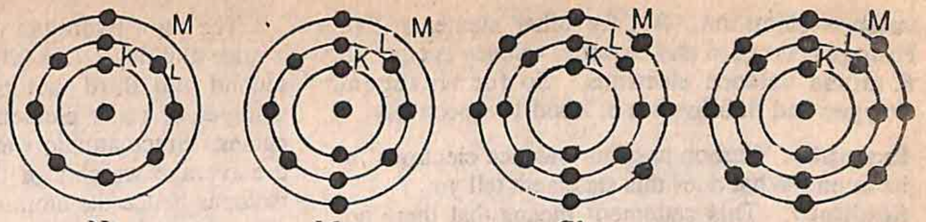








		H	He	Li	Be	Ne
ELEMENT		H	He	Li	Be	Ne
ATOMIC NUMBER		1	2	3	4	10
DISTRIBUTION OF ELECTRONS IN SHELLS	K	1	2	2	2	2
	L			1	2	8



ELEMENT		Na	Mg	Cl	A
ATOMIC NUMBER		11	12	17	18
DISTRIBUTION OF ELECTRONS IN SHELLS	K	2	2	2	2
	L	8	8	8	8
	M	1	2	7	8

Fig. 2.12 Distribution of electrons in various levels

Example : Write the electronic configuration of an element 'X', whose atomic number is 13.

Solution : Atomic Number
= Number of Protons
= Number of Electrons

So No. of Electron = 13

Maximum number of electrons that can be present in the first energy level *i.e.,* K. shell = 2

Maximum number of electrons that can be present in the second energy level *i.e.,* 2 shell = 8

Hence 2 electrons occupy the 1st energy level and 8 of the remaining 11 electrons occupy the 2nd energy level. The remaining 3 electrons will occupy the 3rd energy level.

Hence Electronic Configuration of X =2, 8, 3.

Example : If an element has atomic weight 24 and atomic number 12, how many neutrons does its atom contain ? How many electrons will be present in K, L and M energy levels of its atoms ?

Solution: Atomic weight *i.e.,* Mass Number= 24
Atomic Number = 12

Mass No. = Number of p + Number of n
Atomic Number = No. of protons
= No. of electrons
So No. of protons = No. of electrons = 12
No. of neutrons = Mass No. – Atomic No.
= 24 – 12 = 12

Maximum No. of electrons that can be present in the K energy level = 2
Maximum No. of electrons that can be present in the L energy level = 8

Still there are 2 more electrons which will occupy M energy level

Hence K = 2
L = 8
M = 2

2.7 VALENCE ELECTRONS

From the electronic configuration of the atoms of the elements we can conclude that with the increase in the atomic number, the number of electron in the outermost energy level increases from 1 to 8. When the outermost energy level of an atom is full or has 8 electrons in it then the atom is chemically inert. *For example,* Helium has its first energy level complete, Neon has both its first and second energy levels complete, Argon has 8 electrons in the outermost energy level. Helium, Neon and Argon are all chemically inert. The inertness of these elements is due to the fact that it is very difficult to add electrons to or remove electrons from the outermost energy level of the atoms of these elements. However in the case of other elements the number of electrons in the outermost energy level is less than 8. As a result it is easier to remove electron/electrons from or add to the atoms of these elements. *These electrons present in the outermost energy level of an atom decide the combining capacity of an atom of an element and are called valence electrons.* So valence electrons are the electrons present in the outermost energy level. The energy level containing the valence electrons is called *valence energy level.* It is these electrons that take part in chemical combination.

The valency of an element is given either by its number of valence electrons or 8 minus valence electrons. *For example,* Valency of atoms like sodium, magnesium and aluminium is 1, 2, 3 respectively. It is the same as the number of valence electrons. But for other elements, like Nitrogen, Oxygen and fluorine, valency is equal to 8 minus valence electrons. So for Nitrogen, Oxygen and fluorine it is 3, 2 and 1 respectively.

Example : 'Carbon has four valence electrons' in its atom. What does this statement tell you ?
Solution : This statement means that there are four electrons in the outermost energy level of carbon atom.

Example : Write electronic configuration of an element with atomic number 20. Also write the valence of this element. Is it equal to valence electrons or 8 minus valence electron ?

Solution : Electronic Configuration 2, 8, 8, 2
So number of valence electrons is **two**.
Valency of this elements is also *two* and it is equal to *valence electrons.*

Isotopes : The atoms of a number of elements have nuclei which contain the same number of protons but different number of neutrons. Such atoms whose nuclei have the same value of mass number but different values of atomic number are known as *Isotopes* of the elements. Isotopes were discovered by F.W. Aston in 1919. The isotopes were discovered by measuring e/m ratio with the help of an instrument called Mass spectrograph. **For example :** Hydrogen has the following three isotopes :
${}^{1}_{1}H$; ${}^{2}_{1}D$; ${}^{3}_{1}T$.

The term hydrogen (H) is used for the first isotope and deuterium (D) and tritium (T) refer to second and third isotopes respectively. The isotopes of other elements do not have special names. Since atomic weight of the elements are the average weights of the mixture of different isotopes hence the atomic weight of a number of elements are in fractions.

For example : Chlorine occurs in nature in two isotopic forms with atomic masses 35 and 37. Since they are present in the ratio 3 : 1 the average atomic mass of naturally occurring chlorine is found to be 35·5.

$$= 35 \times \frac{75}{100} + 37 \times \frac{25}{100}$$

$$= \frac{2625}{100} + \frac{925}{100}$$

$$= \frac{3550}{100}$$

$$= 35{\cdot}50$$

Characteristic of Isotopes :

1. All atoms of different isotopes of any element have the same atomic number.
2. They have similar chemical properties because they contain same number of electrons.
3. The atoms of different isotopes have different number of neutrons.
4. These isotopes differ with respect to density, melting point and boiling point because of different masses.

Table 2.3 Composition of some common isotopes

Isotopes	*Notation*	*Atomic Number*	*Mass Number*	*Number of protons*	*Number of neutrons*	*Atomic weight*	*Number of electrons*
Hydrogen-1 (Protinum)	${}^{1}_{1}H$	1	1	1	0	1·008	1
Hydrogen-2 (Deuterium)	${}^{2}_{1}H$	1	2	1	1	2·014	1
Hydrogen-3 (Tritium)	${}^{3}_{1}H$	1	3	1	2	3·022	3
Lithium-6	${}^{6}_{3}Li$	3	6	3	3	6·015	3
Lithium -7	${}^{7}_{3}Li$	3	7	3	4	7·016	3
Carbon-12	${}^{12}_{6}C$	6	12	6	6	12·000	6
Carbon-13	${}^{13}_{6}C$	6	13	6	7	13·003	6
Carbon-14	${}^{14}_{6}C$	6	14	6	8	14·003	6

Isotopes	*Notation*	*Atomic Number*	*Mass Number*	*Number of protouns*	*Number of neutrons*	*Atomic weight*	*Number of electrons*
Nitrogen-14	$^{14}_{7}N$	7	14	7	7	14·003	7
Nitrogen-15	$^{15}_{7}N$	7	15	7	8	15·000	7
Oxygen-16	$^{16}_{8}O$	8	16	8	8	15·995	8
Oxygen-17	$^{17}_{8}O$	8	17	8	9	16·999	8
Oxygen-18	$^{18}_{8}O$	8	18	8	10	17·999	8
Chlorine-35	$^{35}_{17}Cl$	17	35	17	18	34·970	17
Chlorine-37	$^{37}_{17}Cl$	17	37	17	20	36·970	17

X-rays: Wilhelm Konard Roentgen while studying the passage of electricity through gases found that when cathode rays are suddenly stopped by some obstructions such as glass of the discharge tube or a heavy metal (like Tungsten) a new type of rays are produced. Since he did not know what these rays were so he called them *X-rays*. These rays travel in straight line and are not affected by electric and magnetic field. These rays can pass through opaque material such as black paper and also expose the wrapped photographic plate. These rays cannot pass through solid bones (because of the high calcium and phosphorus contents present in bones) so are used to find the fracture between two parts of a bone. The fracture or crack appears as a line as the rays pass through this part and expose the photographic plate but do not pass through the bone which appears as an opaque image on the X-ray film.

RADIOACTIVITY

In 1896 H. Becquerel placed a uranium salt on a photographic plate wrapped in a black paper and found that the photographic plate had the image of the salt. This was due to the fact that uranium salts gave off some radiations that were responsible for penetrating the thickness of the black paper and exposing the photographic plate. Becquerel called this phenomenon *Radioactivity*. At the same time in Paris, Marie Curie and Pierre Curie studied about Radioactivity and found that a radioactive substance keeps on emitting some type of penetrating radiations. They concluded that radioactivity is a property of the atoms. The nuclei of these atoms change into different nuclei by giving off radiations. The unstable atoms are said to be Radioactive. *So Radioactivity is the spontaneous splitting up of the atoms of certain elements into simpler elements and the simultaneous production of radiations.* Such elements are known as radioactive elements. A few examples of radioactive elements are Uranium, Thorium etc.

2.8 TYPES OF RADIATIONS

It was found that the rays emitted from a radioactive substance were not of the same type. On applying a strong magnetic field across the path of the radiations one type of radiations remains unaffected whereas two other types of radiations are deflected from their path. Each of these three types of radiations cause darkening effect on the photographic plate. These radiations are called alpha (α), Beta (ß) and Gamma (γ) rays.

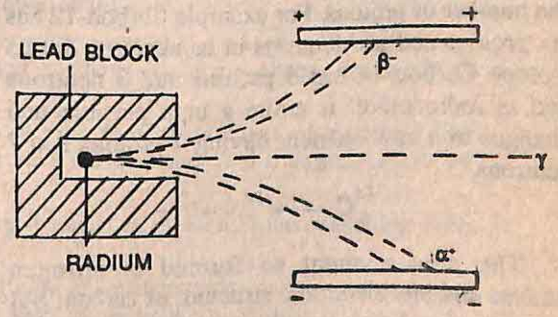


Fig. 2.13 Effect of an Electric field on rays from Radium

Characteristics of Alpha rays :

1. They are positively charged particles.
2. These particles have a mass number of four units and carry two units of positive charge.
3. the symbol is $^{4}_{2}\text{He}$ or $\overset{4}{\underset{2}{\alpha}}$. The superscript represents mass number and subscript represent atomic number.
4. These particles have low penetrating power and travel with one-tenth of the velocity of light.

Characteristics of Beta rays :

1. They are negatively charged particles.
2. They have negligible mass and carry one unit negative charge.
3. The symbol is $_{-1}{}^{\circ}\beta$ or $_{-1}{}^{\circ}e$.
4. They have only a slight penetrating power as compared to gamma rays and travel with nineteenth of the velocity of light.

Characteristics of Gamma rays :

1. They do not carry any charge.
2. They have no mass.
3. The symbol is $^{\circ}\gamma$.
4. They move with the velocity of light and have high penetrating power.

Why is it that certain atoms are radioactive while many others are not ?

In the process of losing the particles there is decrease in the mass and energy of the atom until a stable atom is formed. It is found that heavier atoms are less stable. If we study the nuclear composition of redioactive substances and radioactive isotopes, it is found that an atom becomes unstable when number of neutrons exceed the number of protons. For example Carbon-12 has six protons and six neutrons in its nucleus. But its isotope Carbon-14 has 6 protons and 8 neutrons and is radioactive. It emits a beta particle and changes to a new element having 7 protons and 7 neutrons.

$$^{14}_{6}\text{C} \longrightarrow {}^{14}_{7}\text{N} + {}^{\circ}_{-1}e$$

The new element so formed is nitrogen nucleus and has electronic structure of carbon. But the atom must be electrically neutral so it will pick up one electron from the surroundings and become uncharged nitrogen.

$$^{14}_{7}\text{N} + e^{-} \longrightarrow {}^{14}_{7}\text{N}$$

A question comes to our mind from where does β particle (which is an electron) comes during β emission. It is assumed that one of the neutron changes to proton and electron is emitted.

$$n \longrightarrow p^{+} + e^{-}$$

Radioisotope dating : The radioactive isotopes have the same chemical properties as the normal atoms. This is because the two have the same electronic configuration. Radioactivity is not lost when the radioactive isotope form compounds. So radioactive isotopes finds a number of applications in medicine, biological fields, metallurgy, agriculture and industry. It is also used for estimating the age of fossils. *The method of estimating the age of plants or fossils from the art of its radioactivity is known as Radioisotope dating*. Carbon-14 is produced on earth by cosmic rays bombardment of nitrogen gas in the upper atmosphere

$$^{14}_{7}\text{N} + {}^{1}_{0}\text{n} \longrightarrow {}^{14}_{6}\text{C} + {}^{1}_{1}\text{H}$$

Carbon - 14 so formed decays by a β emission in the following manner :

$$^{14}_{6}\text{C} \longrightarrow {}^{14}_{7}\text{N} + {}^{\circ}_{-1}e$$

As the cosmic rays are constantly falling upon the atmosphere hence the rate of production of Carbon-14 and the rate of its disintegration is in equilibrium. In other words the concentrate of radioactive Carbon-14 in our atmosphere remains constant. So the percentage of radioactive carbondioxide in the atmosphere is constant. Living organisms use carbondioxide. As long as the organism is living it exchanges carbondioxide with the surroundings and the percentage of radioactive carbon remains constant in all living organisms. As soon as the organism dies the exchange ceases. And the percentage of radioactive carbon in the body of living organism goes on decreasing. After 5800 years it will have only half as much Carbon-14 as living organism. So by knowing the ratio of Carbon-14 and Carbon-12 in the sample whose age is to be determined we can estimate the age of the organism.

SUMMARY

2.1. *Cathode rays :* A stream of electrons that move towards anode in the cathode ray tube cathode rays travel in straight line and possess kinetic energy.

2.2. *Electron* : It is a fundamental subatomic particle of which all matter is composed. It carries a unit negative charge and has negligible mass.

2.3. *Positive rays* : A stream of positive ions that move towards cathode in the cathode ray tube. These rays travel in straight and possess kinetic energy.

2.4. *Protons :* It is a fundamental subatomic particle of which all matter is composed. It carries a unit positive charge and has mass equal to $1{\cdot}673 \times 10^{-24}$ grams.

2.5. *Neutron :* It is a fundamental particle of which all matter (except hydrogen) is composed. It is an uncharged particle and have mass equal to proton.

2.6. *Nucleus :* It is the positively charged region of an atom and is composed of protons and neutrons. The mass of an atom is due to the mass of the nucleus alone.

2.7. *Rutherford's Model of an atom :* Atom is considered to be made up of two parts :

(*a*) *Nucleus* : where the total mass of atom is concentrated.

(*b*) *Extra nuclear part* containing the electrons.

2.8. *Atomic Number* : Atomic number of an atom of an element is equal to the number of protons present in its nucleus. Atomic number determines the chemical identity of an atom.

2.9. *Mass Number :* The sum of protons and neutrons present in the nucleus of an atom is known as Mass Number. Protons and Neutrons are collectively known as *nucleons*. The mass number is always a positive integer. Not all atoms of an element have the same mass number.

Energy lebel : Electrons are revolving round the nucleus at a certain definite distance from the nucleus in circular perths called *energy level*. It refers to the definite quantities of energy allowed in atoms. An energy level near to the nucleus have the lowest energy.

Electronic configuration : It is the arrangement of electrons in the different energy levels of an atom. The maximum number of electrons in any energy level is equal to $2n^2$.

Valence Electrons : These are the electrons in the outermost energy level of an atom and decide the combining capacity of an atom of an element valency of an atom is equal to the number of valence electrons or valence electrons – 8.

Isotopes : These are the different atoms of the same element having same number of protons *i.e.*, atomic number but different mass number. Isotopes are due to the different number of neutrons present in the nucleus of an atom. Fractional atomic weights of the elements are due to the presence of different number of isotopes.

X-rays : X-rays are electromagnetic waves with very short wavelengths. They are emitted when fast moving electrons are stopped suddenly by some obstruction such as heavy metals.

Radioactivity is the spontaneous emission of rays from the nucleus of an atom. The nuclei of radioactive elements are so unstable that these elements disintegrate atom by atom.

Alpha Rays are a stream of positively charged nuclei of helium atoms travelling with one-tenth of the velocity of light.

Beta Rays : Beta rays are a stream of negatively charged particles called electrons and travel with a velocity equal to nineteenth of the velocity of light.

Gamma rays : These are the packets of energy having no mass and travel with the velocity of light.

Radioisotope dating : It is the method of estimating the age of plants or fossils from the art of its radioactivity.

QUESTIONS

2.1 Which of the following statement is wrong ?

(*a*) Atoms means uncut.

(*b*) Under ordinary conditions gas is good conductor of electricity.

(*c*) Nature of cathode rays differ from gas to gas.

(*d*) Nature of cathode rays is independent of the material.

(*e*) Cathode rays are deflected towards negative electrode.

(*f*) In the case of cathode rays charge to mass ratio is not constant.

(*g*) Cathode rays do not have kinetic energy.

(*h*) Cathode rays travel in straight line.

(*i*) Cathode rays are made up of positively charged particles.

(*j*) The charge on an electron is one coulomb.

(*k*) Anode rays consist of anions.

(*l*) Nature of positive rays depend upon the nature of the gas in the discharge tube.

(*m*) Anode rays do not travel in straight line.

(*n*) Charge to mass ratio of positively charged particles is same for all gases.

(*o*) In an atom number of protons is equal to number of neutrons.

(*p*) Hydrogen atoms contain no neutrons.

(*q*) Fractional atomic mass of the element is due to the presence of different number of neutrons.

2.2 Complete the following sentences :

(*a*) Cathode rays are a beam of fast moving ....

(*b*) The anode rays obtained from hydrogen gas consists of ....

(*c*) The neutron has ....

(*d*) The fundamental particle not present in a hydrogen atom is ....

(*e*) Almost all the mass of an atom is concentrated in a small region of space called ....

(*f*) The number of positive charges on the nucleus is determined by ....

(*g*) The number of protons in an atom of an element is known as ....

(*h*) An atom has atomic mass number 23 and atomic number 11. The atom has .... protons, ... neutrons and .... electrons.

(*i*) The total number of protons and neutrons in the nucleus of an atom is called its ....

(*j*) The maximum number of electrons that can be present in the outermost energy level of an atom is ....

(*k*) Atomic number of Magnesium is 12. So the number of valence electrons in magnesium is ....

(*l*) All the isotopes of an element contain the same number of ....

(*m*) Isotopes of the same element have different number of ....

(*n*) Some boron atoms have mass number 10 and some have mass number 11. These boron atoms are called ....

2.3 Give experimental evidence which showed that :

(*a*) Cathode rays are made up of particles having kinetic energy.

(*b*) Cathode rays are negatively charged.

(*c*) The atoms of all kinds contain the same negative particles.

(*d*) The whole mass of an atom is centred at its nucleus.

(*e*) Atomic nuclei are positively charged.

(*f*) An atom has a lot of empty space within it.

2.4 Match the descriptions in Column B with the terms/statements in Column A :

Column A	*Column B*
1. Mass of an atom	1. Nuclei that differ in their masses
2. Valence electrons	2. Nucleus
3. Isotopes	3. Take part in chemical bonding

4. Electrons	4. number of protons and neutrons in the nucleus
5. Atomic number	5. extra nuclear region
6. Mass number	6. Number of protons in the nucleus

2.5 Put a tick (√) mark against the suitable choice :

(*i*) Which one of the following descriptions of cathode rays is not correct ?

(*a*) move in a straight line.

(*b*) depend upon the nature of gas used in the discharge tube.

(*c*) cast shadow of the solid object placed in their path.

(*d*) are deflected towards the positive pole of the magnetic field.

(*ii*) Isotopes of an element have

(*a*) same number of neutrons. (*b*) same atomic mass.

(*c*) same atomic number. (*d*) same number of protons.

(*iii*) Aluminium has atomic number 13. Its electronic configuration is :

(*a*) 2, 8, 3. (*b*) 2, 9, 2. (*c*) 2, 10, 1. (*d*) 2, 8, 2, 1.

(*iv*) Atomic number of calcium is 20. The number of valence electrons is :

(*a*) 2 (*b*) 3 (*c*) 10 (*d*) 1.

(*v*) Which of the following statements is not correct for Bohr's model of atom ?

(*a*) The nucleus of an atom is situated at its centre.

(*b*) The electrons move in circular orbits.

(*c*) Electrons can jump from one energy level to another.

(*d*) An electron neither loses nor gains energy in jumping from one energy level to another.

(*vi*) Charge to mass ratio is the same in the case of :

(*a*) anode rays, (*b*) cathode rays, (*c*) X-rays, (*d*) all these.

2.6 Which of the following electronic configurations represent a noble gas, oxygen and potassium ?
(*a*) 2, 8, (*b*) 2, 8, 1 (*c*) 2, 6 (*d*) 2, 8.

2.7 Complete the following table :

Element	*Atomic Number*	*Protons*	*Electrons*	*Neutrons*	*Mass Number*
A	17	–	–	18	–
B	–	–	14	14	–
C	–	9	–	–	19
D	1	–	–	–	2
E	5	–	–	6	–

2.8 In the following table are given mass numbers and atomic numbers of some elements labelled A, B, C, D, E and F.

Element	*Mass Number*	*Atomic Number*
A	2	1
B	3	1
C	3	2
D	6	3
E	9	4
F	11	5

(*a*) How many neutrons are present in one atom of B ?

(*b*) Which atoms are isotopes of the same element ?

(*c*) Which atom has two valence electrons ?

(*d*) Which is the atom of a noble gas ?

(*e*) Which of the atoms have the same number of neutrons ?

2.9 Complete the following table :

Element	*Atomic Number*	*Mass Number*	*Number of Protons*	*Neutrons*	*Electrons*
^{35}Cl	17	35	–	–	–
^{37}Cl	–	37	–	–	–
^{235}U	92	–	–	–	–
^{237}U	–	–	–	–	–
^{3}H	–	–	1	–	–
^{16}O	8	–	–	–	–
^{18}O	–	–	–	–	–

2.10 Draw a labelled diagram to show the presence of charge on cathode rays in the discharge tube.

2.11 What are cathode rays ? State one property of cathode rays.

2.12 What were the observations that led to the conclusions that cathode rays are negatively charged particles ?

2.13 In Rutherford's experiment, some of the alpha particles, when bombarded against a gold leaf, were repelled. Give reasons for this observations.

2.14 What are cathode rays and in what respect do they differ from positive rays (Anode rays) ?

2.15 What is a proton ? How does it differ from neutron ?

2.16 Compare an electron, a proton and a neutron in respect of their relative masses, charges and position in the atom.

2.17 What do you understand by the folowing symbols ?

$_{1}^{1}p$, $_{0}^{1}n$ and $_{-1}^{0}e$,

2.18 (*a*) How was it shown that atomic nuclei are positively charged ?

(*b*) State the location of electrons and protons in an atom.

2.19 Define the terms atomic number and mass number of an element.

2.20 How many electrons will be present in K, L and M energy levels of an atom ?

2.21 Why is an atom neutral inspite of the presence of charged particles in it ?

2.22 A neutral atom has electronic configuration 2, 8, 6. Deduce as many of the following as you can from this information only :

(*a*) Atomic number of the element.

(*b*) Atomic mass number of the element.

(*c*) The number of valence electrons.

(*d*) The name of the element.

2.23 If an element M, has mass number 24 and atomic number 12, how many neutrons does its atom contain ?

2.24 Write the electronic configuration of an element X whose atomic number is 12.

2.25 State the number of electrons present in the outermost energy levels of the atoms of the following elements :

(*a*) Neon (*b*) Chlorine (*c*) Potassium (*d*) Calcium

2.26 What is the number of valence electrons in the atom of an element 'X' having atomic number 9 ? Name the valence energy level of this atom.

2.27 What are valence electrons ? Give three examples to illustrate your answer.

2.28 Why is atomic mass of some elements in fractions ?

2.29 Isotopes of elements have the same chemical properties but different physical properties. Give reasons.

2.30 Write the electronic configuration of the atoms of the elements whose atomic numbers are :

2, 8, 11, 13, 17, 19, 20 and 21.

2.31 Composition of the nuclei of two atomic species A and B is given as under

A		*B*	
Protons	6	Proton	6
Neutron	6	Neutron	8

Give the mass numbers of A and B. What is the relation between the two species and which element or elements do they represent ?

2.32 A naturally occurring sample of chlorine contains 75% of chlorine with mass number 35 and 25% of chlorine with mass number 37. What is the relative atomic mass of the sample ?

□□

CHAPTER 3

HOW ELEMENTS ARE CLASSIFIED ?

LEARNING OBJECTIVES :

1. Explain the need for classifying elements.
2. Describe early attempts of classification of elements.
3. State Mendeleev's Periodic law.
4. List the important features of Mendeleev's Periodic table.
5. List the advantages of Mendeleev's Periodic table.
6. Describe the defects in the Mendeleev's classification.
7. State the modern Periodic law.
8. Enumerate the main features of the long form of the periodic table.
9. Define the terms : Group; period; short period; long period; representative elements; transition elements.
10. List the general characteristics of groups and periods.
11. Describe how the following properties vary in a group and period : Valence electrons ; Valency; atomic size ; ionisation energy ; electron affinity ; metallic and non-metallic character.
12. To predict some properties of an element with respect to its position in the periodic table.

INTRODUCTION :

The main purpose of science is to arrange in a systematic manner the information that has been collected by careful experiments and observations of the materials and the living organisms around us. The enormous abundance and diversity of living organism is remarkable. So scientists have classified these living organisms into plants and animals which have been further classified under different heads namely phylum, class, order, family, genus and species.

Biology studies animals and plants *i.e.*, living organisms and chemistry deals with types of atoms *i.e.*, elements. There are 107 elements known at present which differ widely in their physical and chemical properties and also in the nature of the compounds that they form. So it is very difficult to study the properties of all these elements separately. Hence an attempt has been made from time to time to study the interrelationships among the elements and classify the elements in such a way that the study of a large number of elements is reduced to the study of a few groups of elements. Earlier attempts were made to classify the elements on the basis of their properties namely whether it is a metal or a non-metal, its resistance to the action of acids and alkalies, the nature of the compounds that it forms with other elements.

But in early nineteenth century some form of classification of elements began to emerge. In the present chapter we will study about the various attempts that have been made from time to time to classify the elements and also discuss some of the properties of the elements.

3.1 EARLY ATTEMPTS FOR THE CLASSIFICATION OF ELEMENTS :

The system by which elements are now classified is based on the investigations into the atomic structure of various elements. But 100 years ago there was no basis upon which scientists could classify the elements. The only basis upon which they based their classification was the properties of the elements themselves. Lithium, sodium and potassium were grouped together in a family called *Alkali metals*. They are all metals and react with water to produce caustic solution called alkalies. Fluorine, chlorine, bromine and iodine having similar properties were grouped together and given the name *halogens* (halo- sea ; and gen-generate). However such a classification did not accommodate all the elements. This was because some of the elements of the families had not been discovered and also in some cases these resemblance was only slight.

Following the development of atomic theory and the determination of atomic weights, attempts were made to find a relationship between atomic weight and chemical properties. In 1815 *Prout* suggested that *all the elements* were formed by the simple aggregation of hydrogen atoms since atomic weights of all the elements known at that time were whole number multiple of atomic weight of hydrogen. However soon it was found that a number of elements have fractional atomic weights so it was rejected.

3.2 DOBEREINER'S TRIAD

In 1829 Johann. W. Dobereiner examined certain sets of elements to which he gave the name triads and gave his '*Law of Triads*'. According to it, "Chemically similar elements could be arranged in a group of three in which the atomic weight of the middle element is approximately the arithmetic mean of the other two." Some of the triads are listed in the Table 3.1.

Table 3.1

	Triad			*Mean Atomic weight*
Element	Lithium	Sodium	Potassium	$\frac{7+39}{2} = 23$
Atomic weight	7	23	39	
Element	Calcium	Strontium	Barium	$\frac{40+137 \cdot 4}{2} = 88 \cdot 7$
Atomic weight	40	87·6	137·4	
Element	Chlorine	Bromine	Iodine	$\frac{35 \cdot 5+127}{2} = 81 \cdot 25$
Atomic weight	35·5	80	127	
Element	Sulphur	Selenium	Tellurium	$\frac{32 \cdot 2+129 \cdot 2}{2} = 80 \cdot 7$
Atomic weight	32·2	79·3	129·2	

But this law was rejected because only limited number of elements could be grouped under these heads.

3.3 NEWLAND'S LAW OF OCTAVES

In 1865, Newland arranged the elements in the increasing order of their atomic weights and found that the eighth element, starting from the given one, resemble very closely in properties with the first one. Newland called his generalisation. *The law of Octaves*' His scheme is shown below :

7	9	11	12	14	16	19
Li	Be	B	C	N	O	F
23	24	27	28	31	32	35·5
Na	Mg	Al	Si	P	S	Cl
39	40	48	52	55		
K	Ca	Ti	Cr	Mn		

So if we start with lithium and proceed through the elements with the increasing atomic weight the eighth element is sodium. Sodium had similar properties as lithium. After sodium the eighth element was potassium and potassium had similar properties as lithium and sodium. In the same way berrylium, magnesium and calcium had similar properties. So if we proceed with fluorine the eighth element chlorine would have similar properties and it is so. It was found that it was not true for the elements of higher atomic masses.

3.4 LOTHAR MEYER'S GRAPH

In 1869, the German Chemist Lothar Meyer plotted a graph of atomic volume (atomic volume $= \frac{\text{atomic weight}}{\text{density}}$ against the atomic masses of the elements. He obtained a curve having sharp peaks and broad minima. He observed that elements having similar properties occupy corresponding positions on the graph. On the basis of his curve Lothar Meyer included 56 elements in his periodic table and arranged these in groups and sub-groups. The elements like sodium, potassium, rubidium *i.e.* alkali metals having similar properties occur at the peaks in the graph. The halogens *i.e.*, fluorine, chlorine, bromine and iodine occur at the ascending portion of the curve. However another chemist Mendeleev was making similar investigation in Russia. He made his results known earlier than Lothar Meyer 's work.

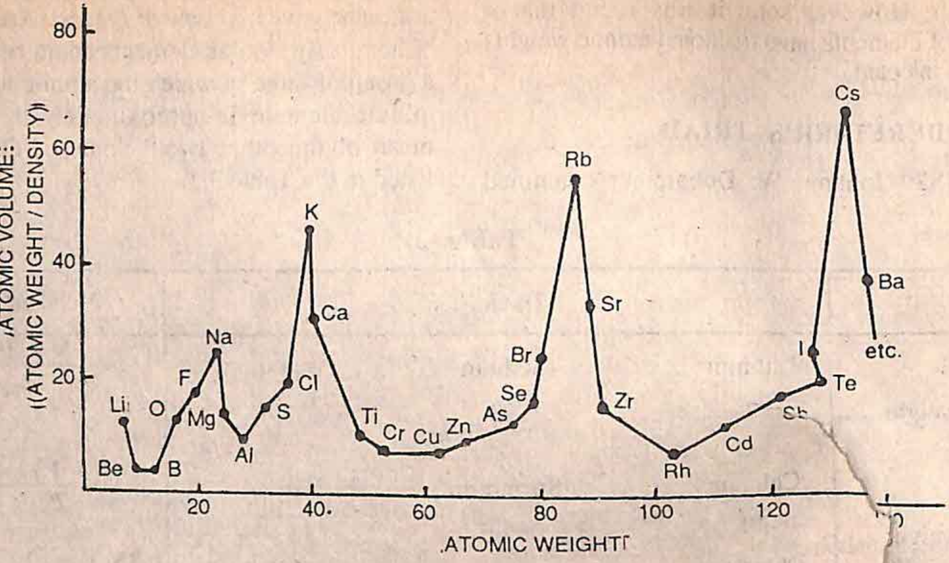


Fig. 3.1. Periodic variation of atomic volume with increasing atomic weigh

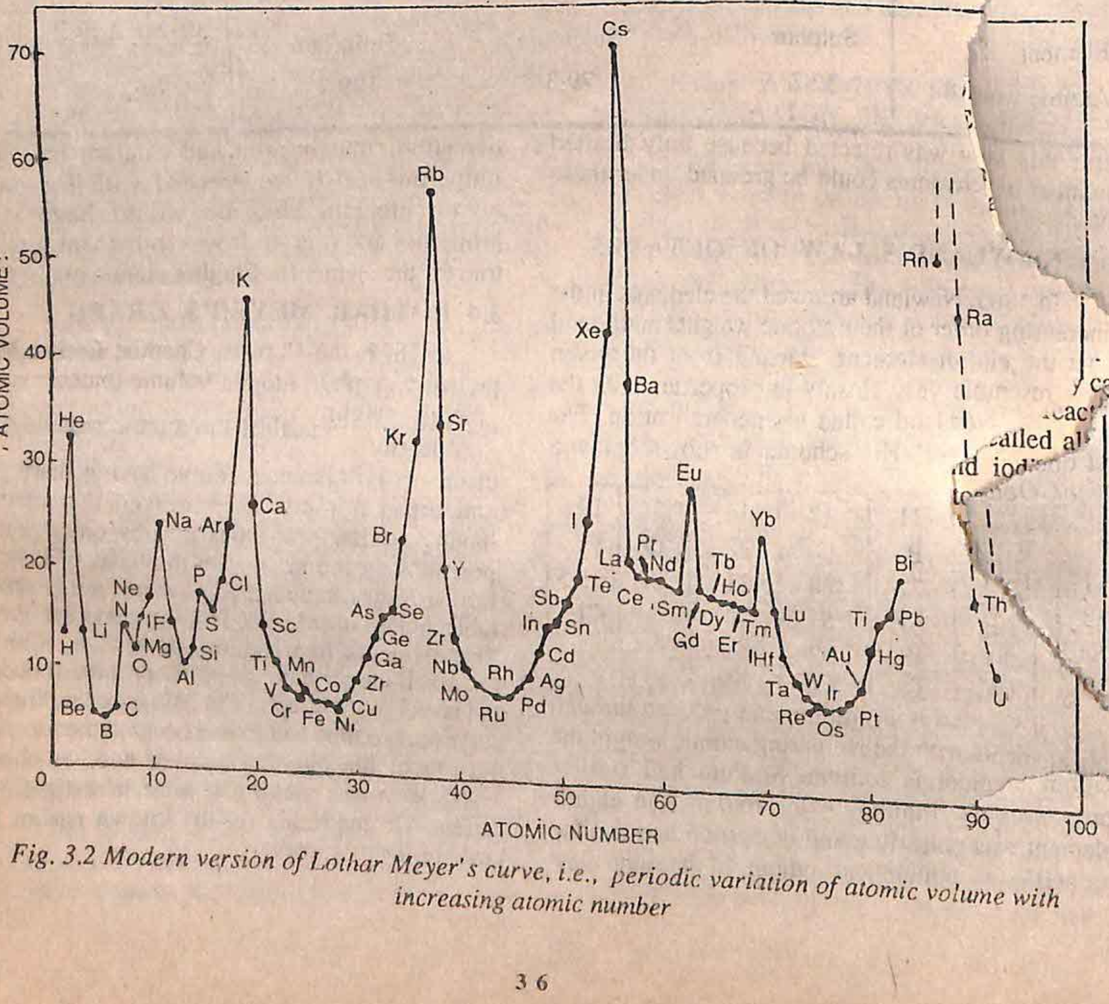


Fig. 3.2 Modern version of Lothar Meyer's curve, i.e., periodic variation of atomic volume with increasing atomic number

3.5 MENDELEEV'S CLASSIFICATION :

Mendeleev developed his ideas about the periodicity of the elements by comparing the chemical properties of the elements and their compounds. He developed further the idea of Newland and arranged the elements in the order of increasing atomic weights in the form of a table known as Periodic Table. The word '*periodic*' means *the recurrence of similar properties of the elements after certain regular interval* :

As the elements are arranged in the tabular form so it is known as table, hence the name Periodic Table. According to Mendeleev's Periodic Law *"Properties of the elements are the periodic function of their atomic masses"*. In other words if the elements are arranged in the order of increasing atomic weights, the elements having similar properties lie one below the other. He arranged the elements in horizontal rows in the order of increasing atomic masses and found that elements having similar properties lie one below the other in the same vertical column. The horizontal rows are called *periods* and vertical columns are known as *groups*. He left gaps for the undiscovered elements and also predicted the properties of some of the undiscovered elements for example. A gap was left between calcium and titanium which he named Eka Boron. It was later discovered and named as *Scandium*. In the same way he named *Eka Silicon* (later on discovered and named as Germanium) and *Eka Aluminium* (later on discovered and named as gallium)

Table 3.2 The periodic table compiled by Mendeleev (1871)

Period	Series	Group I R_2O	Group II RO	Group III R_2O_3	Group IV RH_4 RO_2	Group V RH_3 R_2O_5	Group VI RH_2 RO_3	Group VII RH R_2O_7	Group VIII RO_4
1	1	H=1							
2	2	Li=1	Be=9.4	B=11	C=12	N=14	O=16	F=19	
3	3	Na=23	Mg=24	Al=27.3	Si=28	P=31	S=32	Cl=35·5	
4	4 5	K=39 (Cu=63)	Ca=40 Zn=65	? = 40 ? = 68	Ti=48 ?=72	V=51 As=75	Cr=52 Se=78	Mn=55 Br=80	Fe=56 Co=59 Ni=59 Cu=63
5	6 7	Rb=85 (Ag=108)	Sr-=87 Cd=112	Y= 88 In=113	Zr=90 Sn=118	Nb=94 Sb=122	Mo=96 Te=125	?=100 I=127	Ru=104 Rh=104 Pd=106 Ag=108
6	8 9 10 11	Cs=133 (—) — (Au=199)	Ba=137 — — Hg=200	La=138 — Er=178 Tl=204	?Ce=140 — ?=180 Pb=207	— — Ta=182 Bi=208	— — W=184 —	— — — —	Os = 195 Ir=197 Pt=198 Au= 199
7	12				Tl=231		U=240		

Table 3.2

Table 3.3. A comparison of properties predicted by Mendeleev (1871) with those currently accepted for germanium

Property	*Predicted*	*Current*
Colour	Dirty grey	Greyish-white
Atomic weight	72	72.6
Density (g/cm^3)	5.5	5.35
Heating in air	EsO_2	GeO_2

Property	*Predicted*	*Current*
Properties of oxide :		
Physical state	White solid	White solid
Density (g/cm^3)	4.7	4.703
Compound formed on chlorination	$EsCl_4$	$GeCl_4$
Properties of chloride :		
Boiling point (°C)	below 100	84
Density (g/cm^3)	1.9	1.884
Physical state	Colourless liquid	Colourless liquid
Action of acids	Slight	None by HCl or H_2SO_4,
Action of alkalies	Pronounced	Dissolved by conc. NaOH
Preparation	$EsO_2 + C \rightarrow Es + CO_2$	$GeO_2 + C \rightarrow CO_2 + Ge$

Table. 3.3.

This prediction of the properties of the undiscovered elements reflects the confidence that he had in his Periodic law.

3.5 IMPORTANT FEATURES OF THE MENDELEEV'S PERIODIC TABLE

The main features of his periodic table are :

1. The horizontal rows are known as *series* (now they are called period). There are seven periods numbered from I to VII and are shown in the Table 3.2.
2. The vertical columns are known as *Groups*. There were *eight* groups numbered from I to VIII. However another group comprising of noble gases was discovered later on and was placed in a separate group called *zero group*.
3. Group I to VII are subdivided into subgroups *A* and *B*.
4. As we move along the period from Left to Right there is a regular change from metallic to non-metallic character. For example Sodium is a typical metal and as we move along the period from left to right, chlorine the last member of this period is a non-metal.
5. Group number gives the valency of the elements. With respect to hydrogen the valency increases from 1 to 4 along a group and then decreases to one.

For Example

NaH	MgH_2	AlH_3	S_iH_4	PH_3	H_2S	Hcl
1	2	3	4	3	2	1

And with respect to oxygen the valency increases from 1 to 7

Na_2O	MgO	Al_2O_3	CO_2	P_2O_5	SO_3	U_2O
+ 1	+ 2	+ 3	+ 4	+ 5	+ 6	+ 7

The periodic table has been modified since Mendeleev first formulated it, to give an improved interpretation of the chemistry of the elements. The basic format has been retained. The improved periodic table has been shown in Table 3.3. Each group of the periodic table has been divided into two sub-groups A and B. Noble gases have been placed in Zero group. These gases were not discovered at the time of Mendeleev.

Advantages of Mendeleev's periodic table :

1. It helped in the discovery of new elements. Mendeleev left gaps in his periodic table and also predicted the properties of these elements. For example : Properties of the elements Scanium, Gallium and Germanium were discovered later on were found to exhibit the same properties.
2. Based on the properties of the elements in the periodic table he predicted the errors in the atomic weights of the elements. He corrected atomic weights of gold (Au) and platinum(Pt).
3. He arranged the elements not only in the increasing order of atomic weights but also arranged the elements according to the similarity in their properties. As a result he was able to determine the atomic weights of the elements which were later on confirmed by other methods. *For example* : He placed berryllium in II group so it has valency two.

Atomic Weight = Equivalent weight × Valency

Equivalent weight of Berryllium = 4·5

So atomic weight $= 4.5 \times 2$

$= 9$

This atomic weight was confirmed later on.

Defects in the Mendeleev's Periodic Table : Mendeleev attempted to classify the elements not only on the basis of their atomic weights alone but also based his calssification on the properties of the elements. As a result there were a few defects in his periodic table. These are :

1. *Isotopes could not be provided separate places in the periodic table* : Isotopes are the atoms of the same elements having same atomic number but different atomic weights. *For example :* Oxygen has three isotopes ^{16}O , ^{17}O , ^{18}O so as per Mendeleev oxygen should occupy three positions in the periodic table but it was not so.
2. *Wrong Order of atomic masses of some elements* : While arranging the elements he was guided by two factors namely : (*i*) increasing atomic weight, (*ii*) grouping similar elements. As a result at certain places elements having higher atomic weights were found to occupy position before the elements having lower atomic weights. *For example* : Argon (atomic weight 40) comes before Potassium (atomic weight -39) Tellurium (atomic weight 127·6) comes before iodine (atomic weight 126·9)

So we have seen that all the methods of classifying the elements suffered from one or other defects. Scientists thought, right from Dobernier to Mendeleev everyone considered atomic mass as the fundamental property of an atom (since the only atomic property known at that time was atomic mass) so it is possible that atomic mass may not be the fundamental property of an atom of an element.

3.6 MODERN BASIS OF CLASSIFICATION

Discovery of X-rays and atomic theories gave better understanding of the properties of elements. Moseley carried out a number of experiments and found that the nuclear charge increases by one unit as we move from one element to the next as arranged by Mendeleev in the increasing order of atomic weight in his periodic table. This increase in positive charge on the nucleus was due to the increase in the number of protons. The number of positive charges present in the nucleus of an atom is known as atomic number. Since no two elements can have same atomic number, so he *considered atomic number as the fundamental property of an atom of an element* and arranged the elements in the increasing order of atomic number. Hence the discovery of atomic number brought a major change in the Mendeleev's periodic table.

Modern Periodic Law : Modern Periodic Law was given by Moseley. According to it : "Properties of the elements are the periodic function of their atomic number". So if we arrange the elements in the increasing order of atomic number then elements having similar properties lie one below the other. This is due to the fact that after a certain interval they have similar number of valence electrons hence have similar chemical properties. So it can be seen that the elements which make up a group of the periodic table have the same number of electrons in the outermost energy level. We know all the chemical properties of the elements are decided by the arrangement of electrons in the energy levels. The outermost energy level is particularly important because it contains the valence electrons. These are the electrons that can be given to other atoms or shared with other atoms. So the number of electrons in the outermost energy level (*i.e.*, Valence electrons) are the main factors that determines the sort of chemical reactions the elements will undergo.

While arranging the elements in the increasing order of atomic number the elements that fall into the same group have same outer energy level electrons. In order to have a clear understanding let us consider the elements of IInd horizontal and Ist vertical column.

Looking at the first group we find that lithium, sodium and potassium have one electron each in their outermost energy level after a difference of atomic number 8. This is because sodium has atomic number 8 higher than Lithium and Potassium has atomic number 8 higher than sodium. But after Potassium the next elements are 18, 18 and 32 placed removed from each other. This is because normally electrons fill up their energy levels before progressing on to a new one but in the case of Potassium and after it this does not happen. The reason for this departure from each other you will study in higher classes. However you can see that the number of electrons in the outermost energy level of first vertical column of elements is one. Hence all show similar chemical properties. Hence the reason for the elements having similar properties after a certain interval is due to the same number of valence electrons.

Group

$_1H^1$

PERIOD ⟶ $_3Li$ $_4Be$ $_5B$ $_6C$ $_7N$ $_8O$ $_9F$ $_{10}Ne$

2, 1 2, 2 2, 3 2, 4 2, 5 2, 6 2, 7 2, 8

$_{11}Na$

2, 8, 1

$_{19}K$

2, 8, 8, 1

$_{37}Rb$

2, 8, 18, 8, 1

$_{55}Cs$

2, 8, 18, 18, 8, 1

$_{87}Fr$

2, 8, 18, 32, 18, 8, 1

Table 3.3.(A) A PERIODIC CLASSIFICATION OF THE ELEMENTS THE ATOMIC NUMBER IS GIVEN ABOVE THE SYMBOL OF THE ELEMENT

s-BLOCK

d-BLOCK (TRANSITION ELELMENTS)

p-BLOCK

IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57-71 SEE BELOW	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Pb	82 Tl	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89-103 SEE **	104 Ku+	105 Ha	106 ?												

*LANTHANIDE SERIES	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
** ACTINIDE SERIES	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 Na	103 Lr

for Rt IN DISPUTE

The modified form of the periodic table as shown in Table 3.3A, 3.3B) that is based on atomic numbers removed all the drawbacks that were present in Mendeleev's Periodic Table.

1. *Position of Isotopes* : When the elements are arranged in the increasing order of atomic numbers, the problem of giving different places to isotopes do not arise. Because isotopes are the different atoms of the same element having same atomic number. So they are kept at one place only.
2. It also explained the difficulties arising due to atomic masses in the Mendeleev's table. The anomaly of placing argon before potassium is solved because argon has atomic number lesser than potassium, same is true for tellurium and iodine ; cobalt and nickel.
3. It also explains the periodic recurrence of similar chemical properties of the elements because of periodic recurrence of similar valence shell configuration.

Long form of the Periodic Table : It is constructed by placing the elements in the order of increasing atomic number. In it vertical columns are known as groups. There are 18 vertical columns numbered from 0 to VIII. Group 1 to VII are further subdivided into A and B sub-groups. [But now-a days it is numbered from 1 to 18 starting from one end *i.e.*, left-hand side to the other end. *i.e.*, right-hand side.]

The horizontal rows are known as *periods*. There are seven periods numbered from I to VII. Each period starts with the filling of electrons in the next higher energy level.

First period contains two elements namely *Hydrogen* and Helium. This is because the first period starts with the filling of electrons in the first energy level. First energy level can have at the most two electrons so there are only *two elements* namely hydrogen and helium in the first period.

Second period starts with the filling of electrons in the second energy levels. Second energy level cannot have more than eight electrons, so there are only *eight elements* in the second period from *Lithium* to *Neon.*

Li	Be	B	C	N	O	F	Ne
2, 1	2, 2	2, 3	2, 4	2, 5	2, 6	2, 7	2, 8

Third period starts with the filling of electrons in the third energy level. No doubt third energy level has the capacity of eighteen electrons but as soon as there are eight electrons in the third energy level the next electron enters the fourth energy level. In other words before the completion of 3rd energy level, fourth energy level starts filling up. And with the filling of electrons in the new energy level, new period starts. Hence third period, too, has *eight elements.*

So we have seen that as shown as two electrons enter the fourth energy level, the next differentiating electron starts entering third energy level and when third one is completed *i.e.*, having eighteen electrons then in the next element electrons again starts filling in the fourth energy level till 4th energy level has eight electrons. So in all it is the eighteen electrons that are being filled in the atoms of different elements in this period. Hence there are only *eighteen elements* in the *fourth period.*

Na	Mg	Al	Si	P	O	Cl	Ar
2, 8, 1	2, 8, 2	2, 8, 3	2, 8, 4	2, 8, 5	2, 8, 6	2, 8, 7	2, 8, 8

Such periods *i.e.*, 2nd and 3rd having eight elements each are known as *short periods.*

Fourth period starts with potassium and ends with krypton. It has eighteen elements.

K	Ca	Sc	Ti	V
2, 8, 8, 1	2, 8, 8, 2	2, 8, 9, 2	2, 8, 10, 2	2, 8, 11, 2
Cr	Mn	Fe	Co	Ni
2, 8, 13, 1	2, 8, 13, 2	2, 8, 14, 2	2, 8, 15, 2	2, 8, 16, 2
Cu	Zn	Ga	Ge	As
2, 8, 18, 1	2, 8, 18, 2	2, 8, 18, 3	2, 8, 18, 4	2, 8, 18, 5
Se	Br	Kr		
2, 8, 18, 6	2, 8, 18, 7	2, 8, 18, 8		

Now with the next electron, the new energy level *i.e.*, fifth energy level commences and it leads to the beginning of fifth period.

Fifth Period. It starts with the filling of electrons in the fifth energy level. After two electrons have entered the fifth energy level, next electron starts entering in the fourth energy level till fourth energy level has eighteen electrons and then again fifth energy level is filled till fifth energy level has eight electrons. After it the electrons enters in the sixth energy level and with this the sixth period commences. Hence once again fifth period has eighteen elements. These periods *i.e.*, fourth and fifth periods are known as *long periods.*

If we have a close look at the periodic table then there is a gap of 14 each in VI period from atomic number 57 to 72 and in the VII period from 89 to 104. This is because VI period has 18 elements like IV and V periods but in addition to it there are fourteen elements more due to the filling of electrons in the fourth energy level (which was left incomplete). Hence in all there are thirty-two elements. In the same way there are additional fourteen elements in the VII period which is incomplete at present. All the 14 elements [from atomic nos. 57 to 72 and 89 to 104] have very similar properties hence they cannot be accommodated at one place so are placed in two rows of fourteen elements each just below the periodic table. VI period is known as *very long period.*

A summary of the total number of periods and the elements present in each period is given in the Table :

Periods	*Total Number of Elements*	*From*	*To*
1	2	Hydrogen	Helium
2	8	Lithium	Neon
3	8	Sodium	Argon
4	18	Potassium	Krypton
5	18	Rubidium	Xenon
6	32	Caesium	Radon
7	Incomplete (at present 21)	Francium	–

Note : In the present periodic table we have numbered groups from 1 to 18 according to latest IUPAC system, no doubt old numbering has also been mentioned.

Elements in the same group of the periodic table have same number of valence electrons elements in the groups IA, IIA, IIIA, IVA, VA, VIA, VIIA & Zero are known as *Representative Elements* (In the new periodic table Gp IA & IIA are numbered as Gp I & II, Gp III A to VII A are numbered from 13 to 17 respectively and Group *Zero* is numbered as Group 18.] Elements belonging to Groups III B, IV B, V B, VI B, VII B, VIII, IA & II B are known as *Transition elements.* [In the new periodic table these are numbered from 3 to 12].

Group IA. (Now known as Group I) Elements have one electron in the outermost energy level. These elements are known as *alkali metals* because their hydroxides are strong alkalies. Elements of this group are Li, Na, K, Rb, Cs and Fr.

Group II A. (Now known as group II). Elements belonging to this group have two electrons in the outermost energy level. Their hydroxides are less soluble in water as compared to group IA elements. Hence these are called *alkaline earth metals.*

Group III A (Now known as group XVII) have seven valence electrons. These are known as halogens. (halogen means — sea salt producers) since their salts occur in sea water.

Group Zero (Now known as XVIII group) have eight electrons in the outermost orbit (except helium). These are known as *noble gases* or *inert gases* because these do not react with other elements under normal conditions.

Group IIIB to IIB. (Now known as groups III to XII) Elements belonging to those groups are collectively known as *Transition elements* because their properties are in between elements belonging to group IIA and those belonging to group III A. They are less metallic than IIA group elements but more metallic than the Gp III A elements.

Division of Elements into s, p, d and f blocks?

If you have a look at the periodic table it is as shown on the next page :

TABLE – 3.3 B

Representative elements — d - transition elements — Representative elements — Noble Gases

PERIODIC NUMBER	Group 1	Group 2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	Group 18
1	H 1																	He 2
2	Li 3	Be 4											B 5	C 6	N 7	O 8	F 9	Ne 10
3	Na 11	Mg 12											Al 13	Si 14	P 15	S 16	Ci 17	Ar 18
4	K 19	Ca 20	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30	Ga 31	Ge 32	As 33	Se 34	Br 35	Kr 36
5	Rb 37	Sr 38	Y 39	Zr 40	Nb 41	Mo 42	Tc 43	Ru 44	Rh 45	Pd 46	Ag 47	Cd 48	In 49	Sn 50	Sb 51	Te 52	I 53	Xe 54
6	Cs 55	Ba 56	La* 57	Hf 72	Ta 73	W 74	Re 75	OS 76	Ir 77	Pt 78	Au 79	Hg 80	Tl 81	Pb 82	Bi 83	Po 84	At 85	Rn 86
7	Fr 87	Ra 88	Ac** 89	Unq 104	Unp 105	Unh 106	Uns 107											

f - inner transition elements

* Lanthanides	Ce 58	Pi 59	Nd 60	Pm 61	Sm 62	Eu 63	Gd 64	Tb 65	Dy 66	Ho 67	Er 68	Tm 69	Yb 70	Lu 71
** Actinides	Th 90	Pa 91	U 92	Np 93	Pu 94	Am 95	Cm 96	Bk 97	Cf 98	Es 99	Fm 100	Md 101	No 102	Lr 103

Long form of the Periodic Table of the Elements with their atomic number and ground state electronic configuration. The groups are numbered 1–18 in accord with the 1984 recomendations. This notation replaces the old numbering scheme of 1-A–VIIA, O, 1-B–VIIB and VIII for s, p and d, block elements.

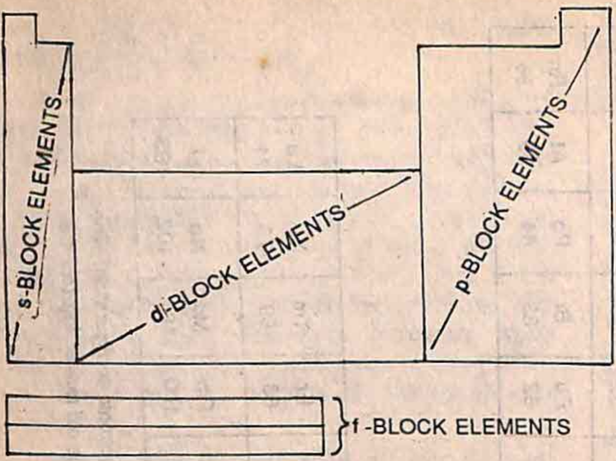


Fig. 3.3

s–Block elements . Elements belonging to I and II groups are known as *s* block elements. Atoms of these elements have incompletely filled outermost energy level and penultimate energy level has eight electrons.

p–Block elements : Elements belonging to groups IIIA to VIII and Zero Gp. [Now numbered from 13 to 18] are known as *p*–block elements. In the atoms of these elements the outermost energy level is in the process of filling and the penultimate energy level is either complete or have 18 electrons in it.

d–Block elements : In between *s* and *p* block elements are a set of elements belonging to Groups IIIB, IVB, VB, VIB, VIIB, VIII (of three rows) IB and XIIB [Now III, IV, V, VI, VII, VIII, IX, X, XI and XII]. These are known as *d*–Block elements. In these cases the penultimate energy level is in the process of filling and the outermost energy level generally contains two electrons.

f–Block elements : These are the elements whose atoms have last three energy levels incomplete. In these cases the anti-penultimate energy level is in the process of filling. The two inner transition series are !

First inner transition series : Cerium through lutetium.

Second inner transition series : Thorium through Lawrencium.

Variation of Properties : **The properties of the elements are determined mainly by their atomic structure. The difference in chemical properties of the elements are mainly due to the differences in the :**

(*a*) Magnitude of the nuclear charge and the number of electrons in the energy levels surrounding the nucleus.

(*b*) The number of energy levels and the number of electrons in these energy levels especially the valence electrons.

(*c*) The distance of the valence electrons from the nucleus.

In the present unit we will discuss about the trends in the properties of the representative elements only *i.e.*, elements belonging to *s* and *p* block of the periodic table.

Valence electrons Group : Elements belonging to the same group of the periodic table have same number of valence electrons. In the case of representative elements, the number of valence electrons is either equal to group number [In the latest P.T. it is equal to Gp.No. (in Gps I & II) or Gp. No. 10 for Gps 13 to 17].

In the case of *d* -Block elements the number of valence electrons is equal to the sum total of the electrons in the outermost energy level plus the number of electrons that exceed 8 in the penultimate energy level. For example in the case of scandium the electronic configuration is 2, 8, 9, 2

Table

Valence electrons for the II period of Representative elements

Group	1/IA	2/IIA	13/IIIA	14/IVA	15/VA	16/VIA	17/VIIA
	Li	Be	B	C	N	O	F
Electronic configuration	2, 1	2,2	2,3	2,4	2, 5	2,6	2,7
Valence electron	1	2	3	4	5	6	7

So Group Number = No. of electrons in the valence energy level plus number of electrons in the penultimate energy level — 8

$= 2 + (9 - 8) = 2 + 1 = 3$

Iron has the configuration 2, 8, 14, 2.

So Group Number = No. of electrons in the valence energy level plus [No. of electrons in the penultimate energy level – 8]

$= 2 + (14 - 8) = 2 + 6 = 8.$

Period : In a period the number of valence electrons goes on increasing from 1 to 8 except helium.

Valency : It is the combining capacity of an atom of an element. The valency of representative elements is equal to the number of valence electrons and/or equal to eight minus the number of valence electrons. Valency of a few elements from the formulae is given in Table.

Trends in Atomic radii of IInd period & I Gp.

Group	I	II	13/IIIA	14/IVA	15/VA	16/VI A	17/VII A	18/0
	Li	Be	B	C	N	O	F	Ne

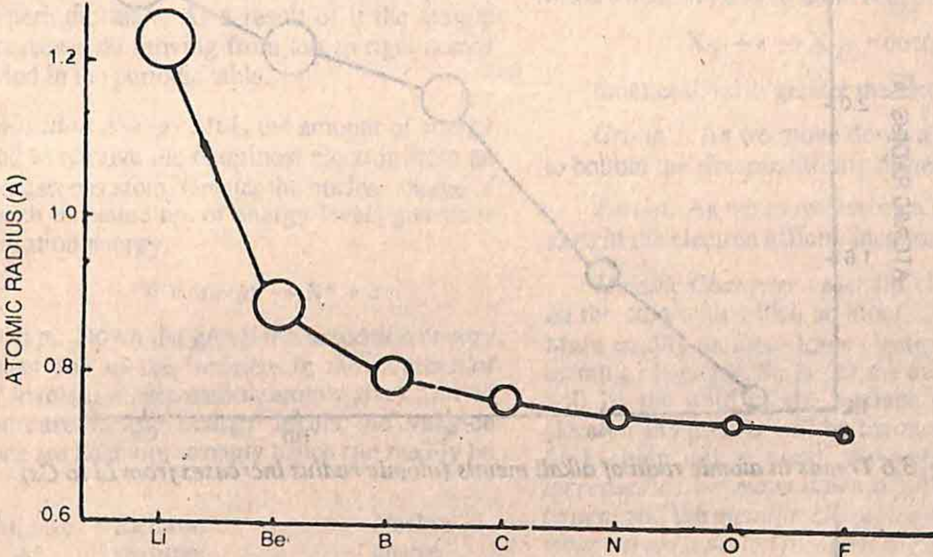


Fig. 3.4 Trends in atomic radii of the 2nd period elements (atomic radius decreases from Li to F)

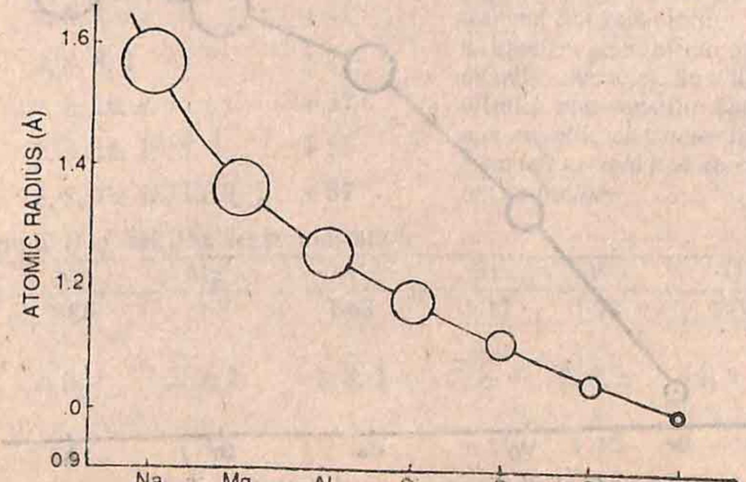


Fig. 3.5 Trends in atomic radii of 3rd period elements (atomic radius decreases from Na to Cl)

Group	1/IA	2/IIA	13/IIIA	14/IVA	15/VA	16/VIA	17/VIIA
	HCl	$BeCl_2$	BCl_3	CCl_4	NCl_3	OF_2	HF
	LiCl	$CaCl_2$	$AlCl_3$	$SiCl_4$	PCl_3	SF_6	HCl
	NaCl	$SrCl_2$	$GaCl_3$	$SnCl_4$	$SbCl_3$		HBr

Atomic size : It is the maximum distance of the outermost electron from the nucleus of an atom of an element *Group* : In a group as we move down a group from top to bottom the atomic size increases. *This is due to the fact that as we move down a group from one element to the other the electrons are being added in the new energy level.* Hence the number of energy levels increases. Although the nuclear charge also increases but the effect due to it is offset by the addition of a new energy level of electrons. Hence the smallest atom of any group will be at the top and the largest at the bottom of the group.

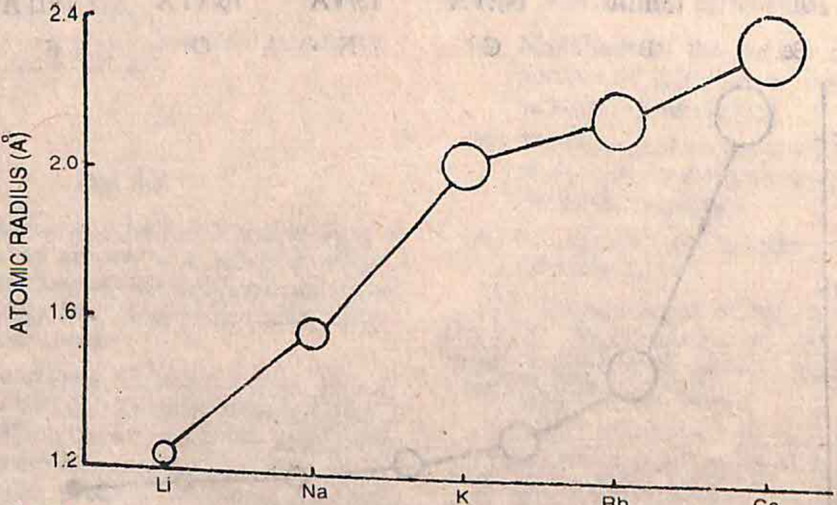


Fig. 3.6 Trends in atomic radii of alkali metals (atomic radius increases from Li to Cs)

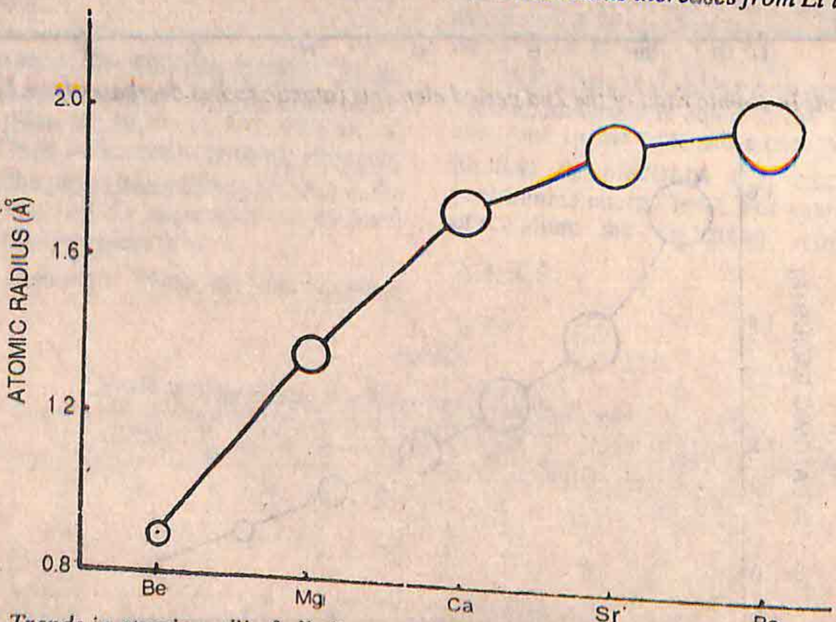


Fig. 3.7 Trends in atomic radii of alkaline earth metals (atomic radius increase from Be to Ba)

Keeping this in mind we can say the smallest atom will be of hydrogen and the largest atom will be of Francium (Last atom of group 1, leaving noble gases).

Period : In general, from left to right across a period, each succeeding element has smaller atomic size than the element preceding it. The atomic size of the noble gases at the end of the periods is the largest because the noble gases have completed outer energy level. Each succeeding element across the period has a nuclear charge of one unit more and also has one more orbital electron than the preceding element although the number of energy levels is same in a given period. Hence the electrostatic forces of attraction increases as we move from one element to the other across a period of the periodic table. As a result of it the atomic size decreases on moving from left to right across the period in the periodic table.

Ionisation Energy : It is the amount of energy required to remove the outermost electron from an isolated gaseous atom. Greater the nuclear charge of atom with the same no. of energy levels greater is the ionisation energy.

$$X + \text{energy} \rightarrow X^+ + e^-$$

Group : Down the group the ionisation energy decreases due to the increase in the number of energy levels (*i.e.*, increase in atomic size). Due to the increase in the energy levels the valence electrons are held less strongly hence can readily be lost.

	At. size A°	Electron structure	Nuclear charge
Li	1·50	2, 1	+ 3
Na	1·86	2, 8, 1	+ 11
K	2·27	2, 8, 8, 1	+ 19
Rb	2·43	2, 8, 18, 8, 1	+ 37
Cs	2·62	2, 8, 18, 18, 8, 1	+ 55
Fr	2·70	2, 8, 18, 32, 18, 8, 1	+ 87

Hence Francium (Fr) has the least ionisation energy.

Period : As we move across a period from left to right the ionisation energy increases. This is because the tendency for each succeeding element to lose electron becomes difficult because the distance of the valence electron from the nucleus is continuously decreasing. So the hold of the nucleus on the valence electrons is continuously increasing.

Electron Affinity : As ionisation energy is a measure of the energy needed to remove an electron from an isolated gaseous atom, electron affinity is the measure of the energy released when an extra electron is added to an isolated gaseous atom. So electron affinity is the amount of energy released when a neutral gaseous atom accepts an electron.

$$X_{(g)} + e \rightarrow X_{(g)} + \text{energy}$$

Smaller the size greater the electron affinity.

Group : As we move down a group from top to bottom the electron affinity decreases.

Period: As we move across a period from left to right the electron affinity increases.

Metallic Character : Metallic character depends on the ease with which an atom can lose electron. More readily an atom loses electron greater is the metallic character. So larger the atomic size lesser will be the hold of the nucleus on the valence electron and greater will be the metallic character. So keeping this in mind, *the metallic character increases as we move down a group from top to bottom* and the *metallic character decreases as we move across a period from left to right.*

Non-metallic character : Non metallic character depends on the ease with which an atom of an element can gain electron. More readily an atom of an element gains electrons, greater will be the non-metallic character. So with the increase in electron affinity non–metallic character increases. Hence *non-metallic character increases across a period from left to right and decreases down a group from top to bottom.*

	Na	Mg	Al	Si	P	O	Cl	Ar
Atomic size(A°)	1·86	1·60	1·48	1·17	1·08	1·06	0·97	1·91
Electronic structure	2, 8, 1	2, 8, 2	2, 8, 3	2, 8, 4	2, 8, 5	2, 8, 6	2, 8, 7	2, 8, 8
Nuclear charge	+ 11	+ 12	+ 13	+ 14	+ 15	+ 16	+ 17	+ 18

Nature of Oxides : Metallic oxides are basic oxides and non-metallic oxides are acidic oxides. Hence as we move across a period from left to right the nature of oxides changes from basic to acidic and on moving down a group the nature of oxides decreases from acidic to basic (if basic then the basic nature of oxides increases)

Some Properties of the Representative elements of Period 3.

Group	1/IA	2/IIA	13/IIIA	14/IVA	15/VA	16/VI A	17/VIIA	18/O
Element	Na	Mg	Al	Si	P	S	Cl	Ar
Melting point(°C)	98	650	660	1410	44	119	− 101	− 189
Electronic configuration	2, 8, 1	2, 8, 2	2, 8, 3	2, 8, 4	2, 8, 5	2, 8, 6	2, 8, 7	2, 8, 8
Atomic radius(*nm*)	0·186	0·160	0·143	0·117	0·110	0·104	0·099	0·154
Nature of Element	Metal	Metal	Metal	Metalloid	Non-metal	Non-metal	Non-metal	Non-metal
Type of Structure	Giant Metallic			Giant molecular	Simple molecular			
Nature of Bonding	Ionic	Ionic	Ionic Covalent	Covalent	P_4 Covalent	S_8 Covalent Ionic	Cl_2 Covalent Ionic	Ar –
Nature of oxides	Basic	Basic	Basic acidic	Acidic Basic	Acidic	Acidic	Acidic	–

Physical Properties such as :–

Melting Point and *Boiling Point ; Density ; Chemical reactivity* : The periodicity of physical properties is related to a periodicity in the type of the element and structure. The melting point first increases and then decreases. Same is true for boiling point as is clear down the above table.

Across a period the elements change from metals through metalloids to non-metals. In period 3, sodium is a very reactive metal, whereas chlorine on the extreme right is a very reactive non-metal. In between the elements show a gradual transition from metals to non-metals. The structure in the elements varies from metallic, through giant molecular, structure in the metalloids to simple molecular structure in the non-metals.

How to predict some properties of an atom ?

In the periodic table the elements have been arranged in the increasing order of atomic numbers. So the position that an element occupies in the periodic table also defines its electronic configuration. The number of valence electrons that an atom of an element has tells us about its group number. Because group number is equal to the number of valence electrons [Now in the new P.T. it is either equal to valence electrons or 10 + valence electrons] OR the group in which the element lies, tells us the number of valence electrons.

This number also helps to predict a number of properties such as valency, metallic character or non-metallic character, size of the atom and so on. It also tells about the reactivity of that element with respect to other elements *e.g.*, A metal is more reactive than the one lying above it in the same group and is also more metallic than the one lying in the next group *w.r.t.* it in the same period. *e.g.*, Have a look at the periodic table. Consider Na that is placed in the Group 1 and period 3 of the periodic table. Its atomic number is 11 so its electronic configurations is $\begin{matrix} K & L & M \\ 2 & 8 & 1 \end{matrix}$. Since it has one valence electron so it belongs to Group I of the periodic table. As there are three energy levels K, L and M, so it belongs to third period. It is a metal because it has one valence electron. A metal has generally one, two or three valence electrons. Greater the number of valence electrons lesser the metallic character. Its valency is one and always form a unipositive ions because it has low ionisation energy and can readily lose this valence electron.

Problem: Three elements A, B and C have atomic numbers 3, 11 and 19. Answer the following questions :

(*a*) To which group/groups of the periodic table do they belong ?

(*b*) To which period/periods of the periodic table do they belong ?

(*c*) Which of these is the most metallic ?

(*d*) What will be the valency of each ?

(*e*) What will be the formula of the

(*i*) Chloride of each.

(*ii*) Oxide of each.

(*f*) What will be the nature of their oxides ?

Solution :

Element	A	B	C
Atomic Number	3	11	19
Electron configuration	2, 1	2, 8, 1	2, 8, 8, 1

(*a*) Since all the three have one valence electron in their atoms hence all belong to group 1A of the periodic table.

(*b*) Period of an element is predicted by the no. of energy levels. Because each period starts with the filling of electrons in the new energy levels. Hence

A belongs to II period

B belongs to III period

C belongs to IV period

(*c*) All are metals. But larger the atomic size greater is the metallic character, so C is the most metallic and A is least metallic.

(*d*) Since they have only one valence electron so the valency of each is one and form unipositive cations.

(*e*) Formula of the chlorides of

A is ACl

B is BCl

C is CCl.

and that of oxides is

$A - A_2O$

$B - B_2O$

$C - C_2O$

(*f*) We know metallic oxides are basic oxides so these oxides will be basic in nature.

SUMMARY

3.1 *Periodic Table :* A tabular arrangement of the chemical elements on the basis of similarities in properties and atomic structure.

3.2 Properties of the elements are the periodic functions of their atomic masses.

3.3 Dobernier classified the chemically similar elements into groups of three and gave his *Law of Triads.* But could not become successful because all elements could not be arranged in triads.

3.4 Newland showed that if elements are arranged in the order of increasing atomic masses, there is repitition of similar properties in every eighth element. But it failed after Manganese *i.e.*, for elements of higher atomic masses.

3.5 In 1869 Lother Meyer plotted a graph of atomic volume versus atomic masses of the elements and observed elements exhibiting similar properties occupy corresponding positions on the graph.

3.6 *Mendeleev's Periodic Law* : The properties of the elements are the periodic function of their atomic masses. He arranged the elements in horizontal rows in the order of increasing atomic masses.

3.7 *Group :* A vertical column in the periodic table is called a group.

3.8 *Period :* A horizontal row in the periodic table is called a period.

3.9 Mendeleev's classification was abandoned due to the fact : (*i*) Isotopes could not be provided separate places.

(*i*) Some elements having higher atomic masses preceded those having low atomic masses.

3.10 *Periodic Law :* The properties of the elements are the periodic function of their atomic numbers. On arranging the elements in the increasing order of atomic number, their properties recur at regular intervals.

3.11 The long form of the periodic table contains 18 vertical columns numbered from 0 to VIII ; Gps I to VII subdivided into A and B and group VIII consisted of three rows but recently it has been numbered from 1 to 18 [Now it is numbered from 1 to 18] and seven horizontal rows called period.

3.12 Elements with atomic numbers 58 to 71 and with atomic numbers 90 to 103 have very similar properties and are placed in two rows of fourteen each just below the periodic table.

3.13 *Representative elements :* The elements in groups IA to IIIA and Zero (Now in groups 1, 2 and from 13 to 18) are known as *representative elements.* Their properties adhere closely to the regular progression on which the periodic table is based.

3.14 *Transition elements :* These are the elements belonging to groups IIIB to IIB (Now numbered from 3 to 12). Their properties do not follow the regular progression of the periodic table.

3.15 *Alkali metals :* Elements belonging to group IA are known as alkali metals. Their hydroxides are strong alkalies.

3.16 *Alkaline earth metals :* Elements belonging to group IA of the periodic table are known as alkaline earth metals. Their hydroxides are less soluble in water.

3.17 *Halogens* : They are non-metals having seven electrons in the outermost energy level and belong to Gp VII A (Now Gp VII) of the periodic table.

3.18 *Noble gases :* They belong to group zero (18) of the periodic table and have eight electrons in the outermost energy level except helium.

3.19 The periodic table is divided into four blocks namely *s, p, d* and *f* blocks.

3.20 *Valence electrons :* In a group the no. of valence electrons remains the same but across a period it increases from 1 to 8.

3.21 *Atomic size :* It is the maximum distance of the outermost electron from the nucleus of an atom of an element. Down a group from top to bottom the atomic size increases but across a period from left to right the atomic size decreases.

3.22 *Ionisation energy :* It is the amount of energy required to remove the outermost electron from an isolated gaseous atom. Down the group the ionisation energy decreases but across a period the ionisation energy increases from left to right.

3.23 *Electron affinity :* It is the measure of the energy released when an extra electron is added to an isolated gaseous atom. Down a group of an atom, electron affinity decreases but across a period it increases from left to right.

3.24 *Metallic character :* It decreases across a period from left to right but increases down a group from top to bottom.

QUESTIONS

3.1 Fill in the blanks :

(*a*) The law of triads was given by ....

(*b*) Mendeleev arranged the elements in the increasing order of ....

(*c*) According to Modern Periodic Law the elements are arranged in the periodic table in the order of their increasing..........................

(*d*) In the periodic table, vertical columns are called .... and horizontal rows are called ....

(*e*) Elements of the same group have .... properties.

(*f*) The properties of an elements are a periodic function of its ....

(*g*) In the long form of the periodic table, there are .... groups and .... periods.

(*h*) The elements which possess.......................... electronic configuration show similar properties.

(*i*) Elements with eight electrons in their outermost energy level are called ....

(*j*) The elements can be divided into .... blocks according to their ....

(*k*) *d-Block* elements are placed between the .... block and the .... block elements.

(*l*) The number of groups present in the *p*–block elements are ....

(*m*) The atomic size decreases in a .... in the periodic table.

(*n*) The amount of energy required to remove the most loosely bound electron from a gaseous atom to form a positive ion is known as ....

(*o*) The amount of energy released when an extra electron is added to an isolated gaseous atom to form a negative ion is known as ....

(*p*) The metallic character .... with increase of atomic number in a group.

(*q*) From left to right within a period, the strength with which the nucleus attracts the electrons in the outer energy level ....

3.2 Put a tick (√) mark against the most appropriate choice :

(*a*) The Law of Octaves was given by (*i*) Lother Meyer (*ii*) Mendeleev (*iii*) Dobernier (*iv*) Newlands.

(*b*) The basis of Mendeleev's classification is :

(*i*) atomic number (*ii*) atomic weight

(*iii*) atomic size (*iv*) atomic volume.

(*c*) The *s*-block elements are :

(*i*) non-metals (*ii*) metals (*iii*) metalloids (*iv*) gases

(*d*) The total number of blocks present in the long form of the periodic table is :

(*i*) 1 (*ii*) 2 (*iii*) 3 (*iv*) 4

(*e*) Which of the following sets of elements do not belong to the same group of the periodic table ?

(*i*) Cu, Ag, Au , (*ii*) Li, Na K, (*iii*) F, Cl, Br, (*iv*) Al, Si, P.

(*f*) The element having the lowest ionisation energy is

(*i*) Li (*ii*) Be (*iii*) B (*iv*) C

(*g*) The atomic size in a period

(*i*) remains the same (*ii*) decreases (*iii*) increases (*iv*) first increases and then decreases

(*h*) The element having lowest electron affinity is

(*i*) Chlorine (*ii*) bromine (*iii*) iodine (*iv*) fluorine.

(*i*) Which of the following elements has the largest atomic size ?

(*i*) Lithium (*ii*) sodium (*iii*) rubidium (*iv*) potassium.

(j) The halogens are present in the .... of the periodic table :

(i) Same group (ii) adjacent groups (iii) same period (iv) s-block

(k) The number of elements in the 4th period of the periodic table are :

(i) 2 (ii) 8 (iii) 18 (iv) 32.

(l) The energy required to remove an electron from an isolated gaseous atom is known as

(i) Ionisation energy (ii) electron affinity
(iii) electronegativity (iv) sublimation energy

(m) Which group of the elements does not belong to the same period ?

(i) C, N, Ne (ii) F, Cl, Br (iii) N, O, F (iv) O, F and Ne

(n) The elements C, N, O and F belong to

(i) same period (ii) same group (iii) same period and also to the same group
(iv) same period but to different groups.

3.3 Point out the correct statements of the following :

(a) Lithium, sodium and potassium were grouped together in a family called alkali metals.
(b) Prout suggested that all the elements were formed by the simple aggregation of hydrogen atoms.
(c) The law of octaves was given by prout.
(d) Properties of the elements are the periodic function of their atomic size.
(e) Mendeleev not only arranged the elements in the increasing order of atomic weights but also arranged the elements according to the similarity in their properties.
(f) Atomic weight = Equivalent weight × Valency
(g) Moseley considered atomic number as the fundamental property of an atom of an element.
(h) Elements having the same atomic masses are called isotopes.
(i) The size of atoms increases in descending a group in the periodic table.
(j) The elements in the centre and the left of the periodic table are known as metals.
(k) Isotopes of the elements have been given separate place in the long form of the periodic table.
(l) Sodium is more metallic than potassium.
(m) Elements with 8 electrons in the outermost energy levels are known as noble gases.

3.4 Give terms used to express the following :

(a) Horizontal rows in the periodic table.
(b) Vertical columns in the periodic table.
(c) Elements having same atomic number but different mass number.
(d) Elements whose atoms contain seven electrons in the valence energy level.
(e) Elements on the extremes right of the periodic table.
(f) Elements having one electron in the outermost energy level.
(g) Elements of second group of the periodic table.

3.5 Match the descriptions under Column B against the statements/terms given under Column A.

Column A	*Column B*
(a) Transition elements	(a) Horizontal rows in the periodic table
(b) Alkali metals	(b) Vertical columns in the periodic table
(c) Halogens	(c) Elements belonging to group VIIA (17) of the periodic table
(d) Alkaline earth elements	(d) Elements belonging to group IA(1) of the periodic table
(e) Groups	(e) Elements belonging to group IIA (2) of the periodic table
(f) Periods	(f) Elements belonging from group IIB to IIB (3 to group 12) of P.T.

(*g*) Inner Transitition elements.

3.6 A, B and C are the elements of a Dobernier's triad. If the atomic mass of A is 7 and that of C is 39 then what is the atomic mass of B ?

3.7 State Newland's Law of Octaves.

3.8 A and B are the two elements having similar properties according to Newland's Law of Octaves. What are the minimum number of elements between A and B ?

3.9 What property did Mendeleev use to classify the elements in his periodic table ?

3.10 State, Mendeleev's periodic law. List two anomalies of Mendeleev's periodic classification of elements.

3.11 Why did Mendeleev leave some gaps in his periodic table ? Name two such elements.

3.12 Why are the noble gases missing from Mendeleev's Periodic table ?

3.13 Where do alkali metals appear on the Lother Meyer Curve ? Do they occupy similar positions ?

3.14 Why was Mendeleev's periodic table more successful than the one proposed by Newland's ? [Hint : Mendeleev left gaps in his periodic table. He also predicted the properties of some of the missing elements].

3.15 State Modern Periodic Law.

3.16 List one anomaly of Mendeleev's periodic classification and how this has been removed in the longform of the periodic table.

3.17 What is the reason for periodicity in the properties of elements ?

3.18 How many periods and groups are there in the long form of the periodic table ?

3.19 What is the basis for the classification of elements for modern periodic table ?

3.20 How do the following properties change on moving across a period ?

(*a*) Valence electrons (*b*) atomic size (*c*) metallic character.

3.21 How do the following properties change on moving down a group ?

(*a*) Valence electrons (*b*) atomic size (*c*) metallic character.

3.22 From the standpoint of atomic structure what determines which element will be the first and which the last in a period of the periodic table ?

3.23 Among the elements of the second period Lithium to Neon pick out the element :

(*a*) with the highest ionisation energy.
(*b*) with the highest electron affinity.
(*c*) with the largest atomic size.
(*d*) that is the most reactive metal.
(*e*) that is the most reactive non-metal.

3.24 Why does the ionisation energy increase as we go across a period from left to right ?

3.25 Why does the atomic size increase as we move down a group from top to bottom ?

3.26 Oxygen (atomic number - 8) and sulphur (atomic number 16) are elements in the same group of the periodic table. Write down their electronic configuration. To which group of the periodic table do they belong ?

3.27 For the main groups of the Representative elements of the periodic table, the metallic properties of the elements vary approximately with their positions as shown in the chart below :

1/IA	2/IIA	13/IIIA	14/IVA	15/VA	16/VIA	17/VIIA	18/0	According to latest system previously
H						He		
A						B		
C						D		

(*a*) Where will you find the most metallic element ?
(*b*) Where will you find the most non-metallic element ?
(*c*) Where will you find the smallest atom ?
(*d*) Which will have the maximum ionisation energy ?
(*e*) Which of these is halogen ?
(*f*) Which of these is an alkali metal ?

3.28 Sodium and aluminium have atomic numbers 11 and 13 respectively. They are separated by one element in the periodic table and have valencies one and three respectively. Fluorine and sodium are also separated by one element in the periodic table (Fluorine has atomic number 9) and yet both have valency of one.Explain.

3.29 State one reason for keeping chlorine and bromine in the same group of the periodic table.

3.30 Given below are the melting points and atomic radii of the elements X, Y and Z of the periodic table each having *n* electrons in the outermost energy level of their atoms :

Elements	Melting Point(°C)	Atomic Radii(Å)
X	214	1·20
Y	121	1·31
Z	101	1·36

Answer the following :

(*a*) Will the valencies of the elements be same or different ?
(*b*) Which element will have the highest atomic number ?
(*c*) Do they belong to same group or period ?

3.31 Three elements A, B and have atomic numbers 2, 10 and 5 respectively. Which of these belong to
(*a*) same period ?
(*b*) same group ?

3.32 Given below is a part of the periodic table :

Li	Be
Na	
K	
Rb	
Cs	
Fr	Ra

As we move vertically downward from Lithium to Francium :

(*a*) what happens to the size of the atom ?

(*b*) what happens to their metallic character ?

3.33 Consider the elements (Li, Be, B, C, N, O, F, Ne) in the second period of the periodic table

(*a*) Which elements

(*i*) form cations ?

(*ii*) form anions ?

(*iii*) form a chloride of the formula ACl_3 where A is an element of the second period?

(*iv*) react together to form a compound of the formula AB. (A and B belong to second period) ?

(*v*) exist as diatomic molecules at room temprature ?

3.34 Predict which ion in each pair is the larger :

(*a*) Mg^{2+} or Ca^{2+} (*b*) K^+ or Ca^{2+} (*c*) F^- or Cl^-

(*d*) Br or Br^{-1} (*e*) Na or Na^+

3.35 Name the four blocks of the elements in the periodic table. State their characteristics.

3.36 Write short notes on the following ?

(*a*) Dobernier's Triad (*b*) Newland's Law of Octaves

(*c*) Ionisation Energy (*d*) Electron Affinity.

3.37 (*a*) Why does the size of the atom progressively become smaller from sodium to chlorine ?

(*b*) Why is chlorine monovalent white silicon is tetravalent ?

(*c*) Why is argon larger in size than chlorine ?

□□

CHAPTER 4

CHEMICAL BONDING

LEARNING OBJECTIVES :

1. Define Chemical bond.
2. Explain the basis for the formation of a chemical bond.
3. Describe the formation electrovalent bond.
4. List the groups of periodic table whose element form electrovalent bond.
5. Describe the formation of covalent bond. (single, double and triple).
6. Compare the properties of compounds containing electrovalent and covalent bonds.
7. Give examples of both kinds.
8. Define electrovalency and covalency.
9. Describe shape of simple covalent molecules as CH_4 ; H_2O ; NH_3 .
10. Define polymers.
11. Define polar covalent bond ; coordinate covalent bond

INTRODUCTION

Let us have a look at the electronic configurations of the atoms and ions of the elements of period III.

Element	Na	Mg	Al	Si	P	S	Cl	Ar
Electronic configuration	2, 8, 1	2, 8, 2	2, 8, 3	2, 8, 4	2, 8, 5	2, 8, 6	2, 8, 7	2, 8, 8
No.of Electrons in Valence Energy level	1	2	3	4	5	6	7	8
Ion Formed	Na^+	Mg^{2+}	Al^{3+}	-	-	S^{2-}	Cl	-
Electronic configuration of ions	2, 8	2, 8	2, 8		-	2, 8, 8	2, 8, 8	-

So we have seen that Sodium, Magnesium and Aluminium lose electrons from their valence energy level and form positively charged ions [Na^+ ; Mg^{2+} ; Al^{3+}] to have octet in its outermost energy level { configuration that of neon}.

Elements like sulphur and chlorine (of groups 16, 17) gains electrons and form negatively charged ions (S^{2-} ; Cl^-) so as to have octet in its outermost energy level (configuration that of argon).

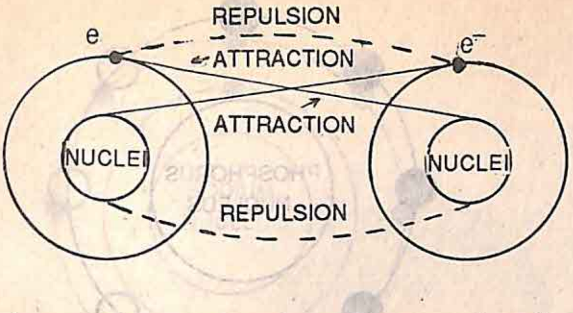
But what about the elements like silicon and phosphorus (of groups 14, and 15).

These elements generally do not form ions. Is it that these elements do not form compounds ? OR If they form compounds is it that their atoms do not tend to have a stable octet in the outermost energy level ? But it has been observed that these elements also form compounds like $SiCl_4$; PCl_3 ; PH_3 etc. It is clear that they cannot attain a stable octet by electron transfer since it will involve the formation of ions. Instead of this these atoms attain the stable configuration by the sharing of electrons with other atoms in their compounds.

In the present chapter we will study all about it *i.e.*, how the atoms combine to form molecules of elements and compounds ? What forces hold atoms together in a molecule ? Why is it that fixed numbers of atoms will join together to form molecules of elements and compounds ? Why cannot any number of atoms join together ?

4.1 CHEMICAL BOND : ITS MEANING

When two similar or dissimilar atoms approach near each other, both forces of attraction and repulsion comes into play. As a result of these forces, the electrons present in the outermost energy levels of the atoms rearrange to attain a



Attractive and repulsive forces that comes into play when two atoms approach near each other.

stable octet. This rearrangement of electrons results in the formation of molecules.

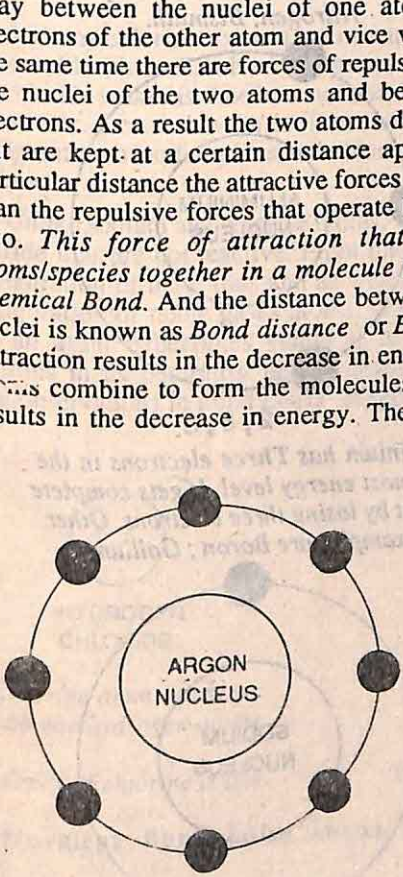
Let us consider the formation of hydrogen molecule. As the atoms of hydrogen approach near each other, a force of attraction comes into play between the nuclei of one atom and the electrons of the other atom and vice versa. But at the same time there are forces of repulsion between the nuclei of the two atoms and between their electrons. As a result the two atoms do not merge but are kept at a certain distance apart. At this particular distance the attractive forces are stronger than the repulsive forces that operate between the two. *This force of attraction that holds the atoms/species together in a molecule is known as chemical Bond*. And the distance between the two nuclei is known as *Bond distance* or *Bond length*. Attraction results in the decrease in energy. Hence atoms combine to form the molecules because it results in the decrease in energy. The amount of energy that is released when atoms combine to form molecules is known as *Bond Energy*. For breaking up a bond that much energy has to be supplied from outside. The amount of energy required to break up the bonds is known as Bond dissociation energy.

4.2 ELECTRONIC CONFIGURATION AND CHEMICAL BONDING

Atoms like Helium, Neon, Argon do not form any compound and are present in the nature as monoatomic gases. These elements belong to group zero (now numbered as 18 group) of the periodic table. The combining capacity of these elements is considered as zero. If we have a look at the electronic configuration of the atoms of these elements, as shown in Table 4.1, it is found that except Helium each has eight valence electrons in the outermost energy level and exist independently in nature. These are known as inert gases because they are totally unreactive, all have full outer energy level of eight electrons in their atoms. Atoms which do not have eight electrons in the outermost energy level tend to combine together, if by doing so they acquire eight electrons in the outermost energy level.

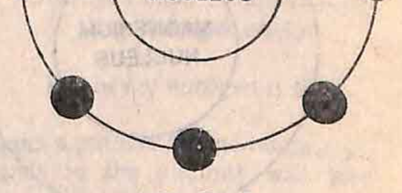
Table 4.1

He	2
Ne	2, 8
Ar	2, 8, 8
Kr	2, 8, 18, 8
Xe	2, 8, 18, 18, 8
Rn	2, 8, 18, 32, 18, 8



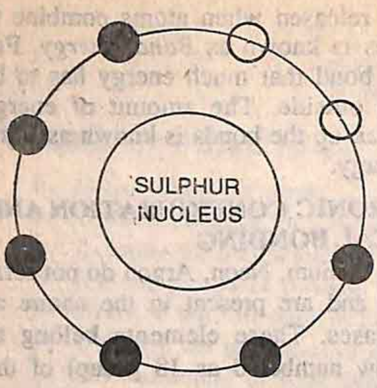
2 8; 8

Argon has Eight electrons in the outermost energy level. It already has a complete octet. Its valency is therefore zero. Other examples are : Neon, Krypton ; Xenon.



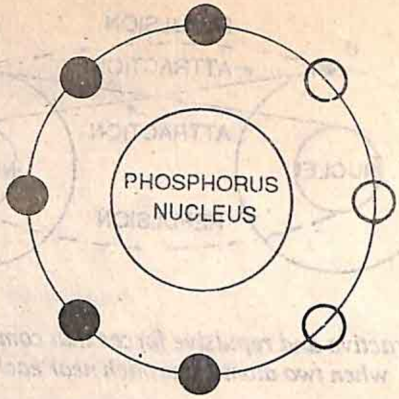
2 8. 7

Chlorine has Seven electrons in the outer most energy level. It gets complete Octet by gaining one electron (shown by dotted line) other examples are Bromine,Iodine.



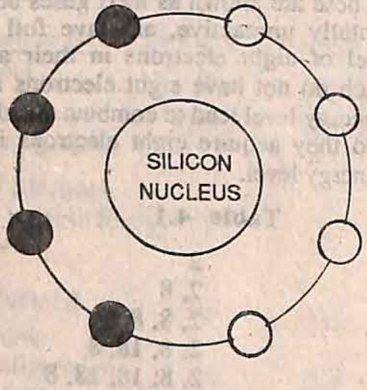
2 , 8 , 6

Sulphur has Six electons in the outermost energy level. It gets complete octet by gaining/ shairing two electrons (shown by dotted line) other examples are Oxygen, Selenium.



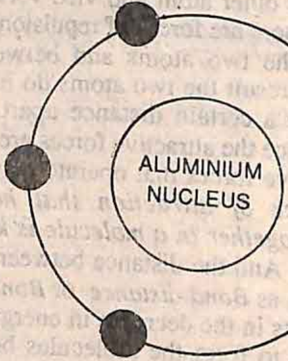
2 , 8 , 5

Phosphorus has Five electrons in the outermost energy level. It gets complete Octet by sharing or gaining three electrons (shown by dotted line) other examples are Nitrogen, Bismuth.



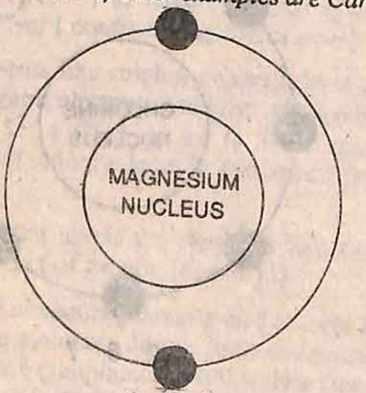
2 , 8 , 4

Silicon has Four electrons in the outermost energy level. It gets complete Octet by sharing four electrons (shown by dotted line) other examples are Carbon, Tin.



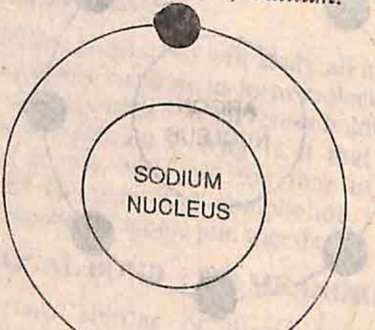
2 , 8 , 3

Aluminium has Three electrons in the outermost energy level. It gets complete Octet by losing three electrons. Other examples are Boron ; Gallium.



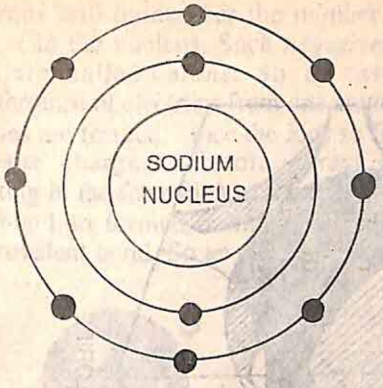
2, 8 , 2

Magnesium has Two electrons in the outermost energy level. It gets a complete Octet by losing two electrons. Other examples are Barium, Strontium, Calcium.



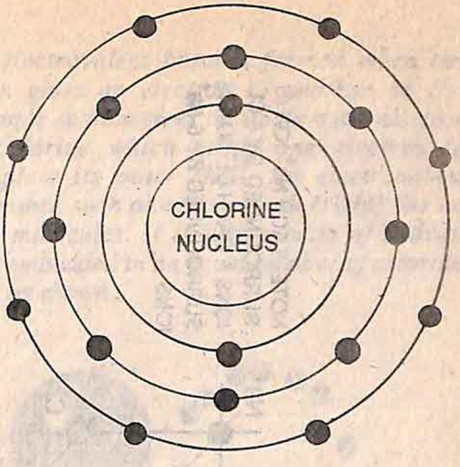
2 . 8 .1

Sodium has One electron in the outermost energy level. It gets a complete Octet by losing one electron. Other examples are Potassium, Lithium.



SODIUM ION

2, 8



CHLORINE ION

2, 8, 8

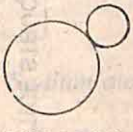
Sodium chloride exists free in nature. When sodium chloride is dissolved in water, sodium ions and chloride ions exist free in water and do not react with water. Let us have a look at the electronic configuration of sodium ions and chloride ions it is as above :

Unlike sodium and chlorine atoms, sodium and chloride ions are not reactive. From the electronic configuration of these ions and as well that of the configuration of noble gases atoms. We can infer that an atom is/becomes stable if it has eight electrons in its outermost energy level. Keeping these observations in mind *Lewis* suggested *Octet rule*. According to it, an atom tends to gain, lose or share electrons so as to have eight electrons in the outermost energy level." In other words an atom tends to have the configuration of a nearest noble gas either by

(*a*) Transferrance of electrons

(*b*) Sharing the electrons.

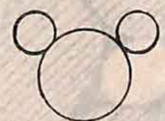
The number of electrons that an atom of an element can lose, gain or share is known as the valency of an atom. It is responsible for the joining power of an atom of an element.



HYDROGEN
CHLORIDE

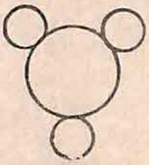
(Chlorine atom combines with one hydrogen atom)

Valency of chlorine is one



WATER

(Oxygen atom combines with two hydrogen atoms)
Valency of Oxygen is two



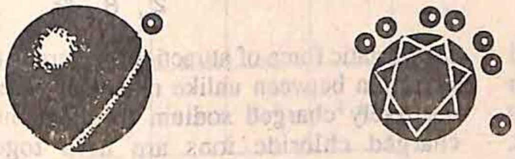
AMMONIA

(Nitrogen atom combines with three hydrogen atoms)

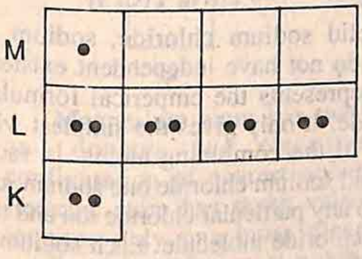
Valency of nitrogen is three

Electrovalent Bond [Also known as Ionic bond]

An ionic bond is formed by the *complete* transferrance of electrons from one atom to another. As a result of the transferrance of electrons, charged particles called ions are formed. Atoms that have lost electrons acquire a positive charge because the number of protons in the nucleus will now outnumber the number of remaining electrons. Such *positively charged ions* are called *cations*. On the other hand the atoms that have gained electrons acquire *negative charge* because now the number of electrons will outnumber the number of protons present in the nucleus. Such *negatively charged* ions are called *anions*. So as result of the transferrance of electrons from one atom to another, the ions are formed. Since the ions so formed have opposite charge, therefore attract each other resulting in the formation of a bond between them. The bond so formed is known as ionic bond or electrovalent bond. So an electrovalent bond or an *ionic bond is one that is formed due to the transferrance of one or more electrons from one atom to another. The number of electrons that an atom loses or gains is known as the electrovalency of that atom :*



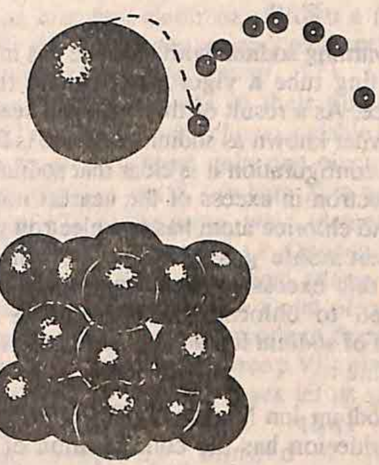
Consider the reaction between sodium and chlorine to form sodium chloride. The electronic structure of sodium and chlorine is as shown in the figure.

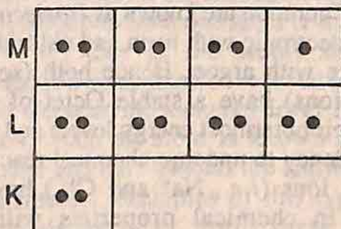


Sodium atom

Electronic Configuration 2, 8, 1

An electrovalent bond is formed when one atom gives an electron to another. In this example sodium gives its single outer electron to chlorine which needs one electron to complete its outer shell. An electrovalent substance such as common salt (right) has no true molecules. It is just a mass of sodium ions embedded in an equal number of chloride ions as shown.

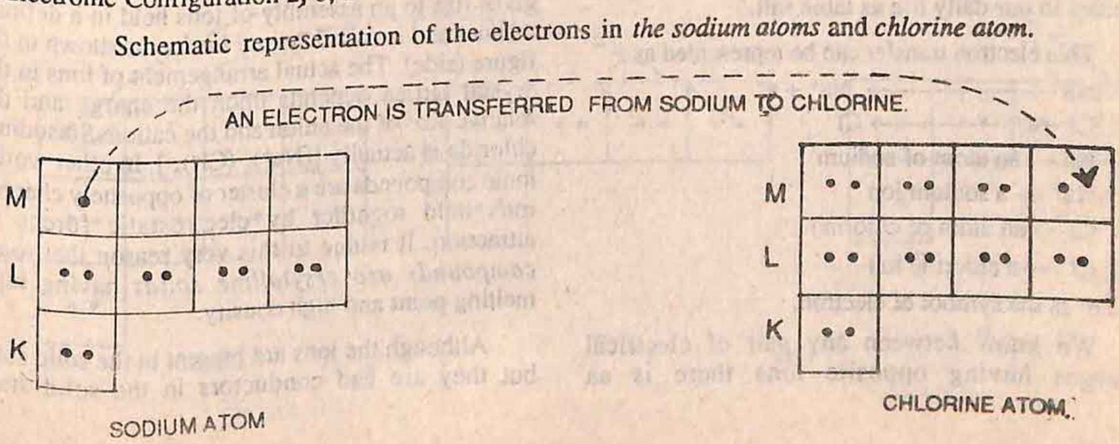


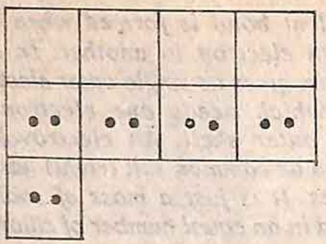


Chlorine atom

2, 8, 7

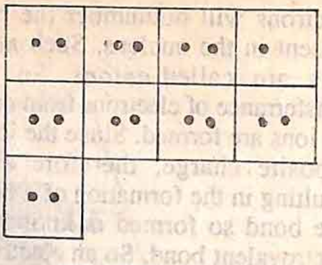
Schematic representation of the electrons in *the sodium atoms* and *chlorine atom.*





SODIUM ION (Na^+)

2, 8



CHLORIDE ION (Cl^-)

2, 8, 8

On warning sodium with chlorine gas in a hard glass boiling tube a vigorous chemical reaction takes place. As a result of this chemical reaction a white powder known as sodium chloride is formed. From the configuration it is clear that sodium atom has an electron in excess of the nearest noble gas (Neon) and chlorine atom has one electron short of the nearest noble gas (Argon). So during the reaction this excess of electron from sodium is transferred to chlorine atom resulting in the formation of sodium ion and chloride ion as shown in the figure :

So sodium ion has the configuration of neon and chloride ion has the configuration of argon. Species (atoms or ions) that have the same electronic configuration are known as *Isoelectronic*. So Na^+ is isoelectronic with neon and chloride ion is isoelectronic with argon. Hence both (sodium and chloride ions) have a stable Octet of eight electrons in their outermost energy levels and show no further tendency to undergo chemical reactions. The resulting ions (*i.e.*, Na^+ and Cl^-) have no resemblance in chemical properties with the original elements (Na and Cl atoms). We started with two very reactive elements – one a metal and other a poisonous gas and now we have a white powder that is neither metallic nor poisonous and is used in our daily life as table salt.

This electron transfer can be represented as :

$Na \longrightarrow Na^+ + e^-$

$Cl + e^- \longrightarrow Cl^-$

Na — an atom of sodium

Na^+ — a sodium ion

Cl — an atom of chlorine

Cl^- —a chloride ion

and e^- is the symbol of electron.

We know between any pair of electrical charges having opposite ions there is an electrostatic force of attraction as there is a force of attraction between unlike magnetic poles. Hence positively charged sodium ions and negatively charged chloride ions are held together by electrostatic attractions. So we can conclude that an *ionic bond is formed due to the attraction between opposite charged ions that are formed by the complete transferrance of electrons from the outer energy level of one atom to another.* So the reaction between sodium and chlorine can be written as

$Na + Cl \longrightarrow Na^+ \; Cl^-$

Since chlorine acts as diatonic molecules *i.e.*, Cl_2. So it can be rewritten as

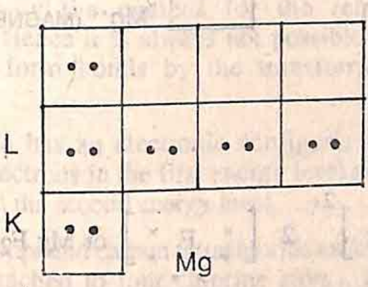
$2Na + Cl_2 \longrightarrow 2[(Na^+)\,(Cl^-)]$

In the solid sodium chloride, sodium and chloride ions do not have independent existence. Hence Nacl represents the emperical formula of sodium chloride. It only gives the simplest whole number ratio of the combining atoms. In fact in the case of solid sodium chloride one sodium ion is not attached to any particular chloride ion and there is no sodium chloride molecule. Each sodium ion has six chloride ions around it at an equal distance from it and in the same way each chloride ion has six sodium ions at an equal distance from it. This gives rise to an assembly of ions held in a definite manner and is called *crystal lattice* as shown in the figure (side). The actual arrangement of ions in the crystal lattice depends upon the charge and the relative size of the anion and the cation. So sodium chloride is actually $[(Na^+)_n\,(Cl^-)_n\,]$. In other words ionic compounds are a cluster of oppositely charged ions held together by electrostatic forces of attraction. It is due to this very reason that *ionic compounds are crsytalline solids* having high melting point and high density.

Although the ions are present in the solid state but they are bad conductors in the solid state because the ions are not free to move and are held at fixed places. On heating, when they melt, they become good conductors because now ions are in a position to move.

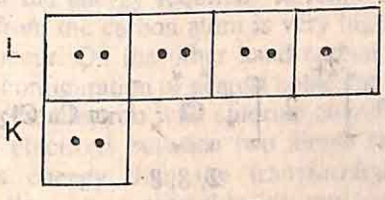
4.3 FORMATION OF MAGNESIUM FLUORIDE COMPOUND

Magnesium atom has the electronic configuration 2, 8, 2.



2, 8, 2

Fluorine has the electronic configuration 2, 7.

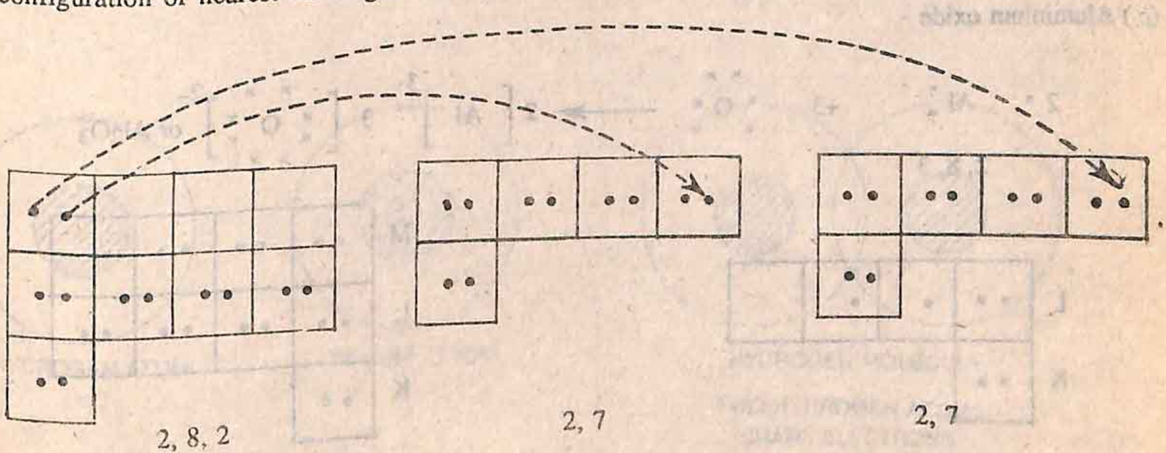


2, 7

Magnesium atom has two valence electrons and donates two electrons to attain the configuration of nearest noble gas atom (neon) Fluorine atom has seven valence electrons so it requires only one more electron to attain the configuration of nearest noble gas atom (neon). Since one magnesium atom donates two electrons so two fluorine atoms approach and gain one electron each.

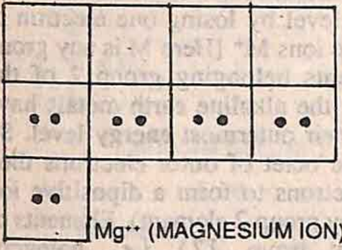
Elements of the group 1 of the periodic table *i.e.*, alkali metals achieve a stable Octet in the outermost energy level by losing one electron to form a unit positive ions M^+ [Here M is any group 1 element] Elements belonging group 2 of the periodic table *i.e.*, the alkaline earth metals have two electrons in their outermost energy level. So to achieve a stable octet of outer electrons they must lose two electrons to form a dipositive ion M^{2+} (Here M is any group 2 element). Elements of group VII A (now group 17) *i.e.*, halogens have seven electrons in their outermost energy level and in order to achieve a stable Octet of outer electrons they must gain one electron to form a unit negative ion X^- (where X is any halogen of group 17). In the same way elements of group VI A (now group 16) *i.e.*,Oxygen family have 6 electrons in their outermost energy level and in order to achieve a stable Octet of outer electrons they must gain two electrons to form a dinegative ion, R^{2-} (R is any group VIA element). After considering the above cases let us see **which are the best atoms that can be electrovalently bonded ?** We know an electrovalent bond formation requires transferrance of electrons from one atom to another. Hence

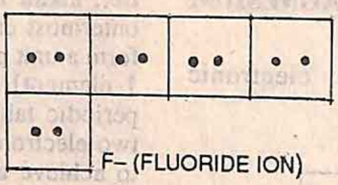
(*i*) One of the atoms should readily lose electron so as to achieve a stable Octet of electrons in the outermost energy level and form a positive ion. Such an atom is known as *electropositive atom* and the elements are known as electropositive elements. Examples of this type are most of the metals. In other words **these atoms have low ionisation energy.**

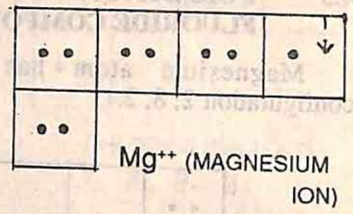


2, 8, 2 2, 7 2, 7

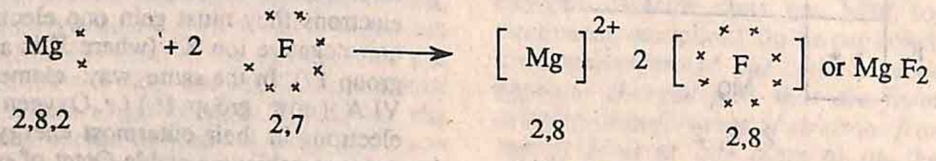
(*ii*) The other atom should have a tendency to readily accept the electrons to achieve a stable Octet of electrons in the outermost energy level and form negatively charged ions. Such atoms are known as *electronegative atoms* and the elements are known as electronegative elements.


Mg⁺⁺ (MAGNESIUM ION)


F– (FLUORIDE ION)

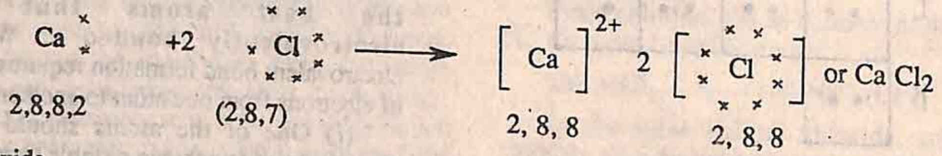

Mg⁺⁺ (MAGNESIUM ION)

The complete reaction can be written as

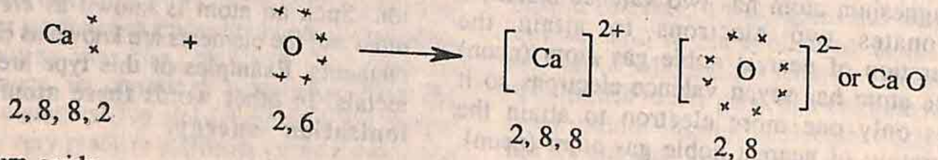


A few more examples are :

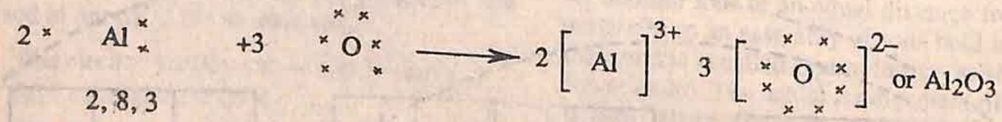
(*a*) Calcium chloride

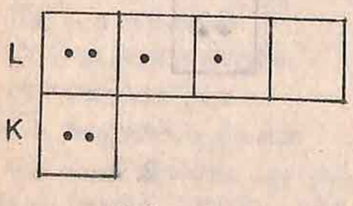


(*b*) Calcium oxide

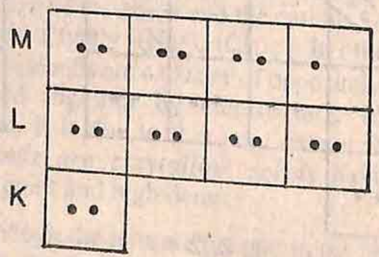


(*c*) Aluminium oxide





Carbon (2, 4)



Chlorine (2, 8, 7)

(*iii*) Thirdly the ions so formed being oppositely charged attract each other to form a crystal lattice. Hence strong forces of attraction should exist between the two.

The Covalent Bond : In the previous discussion we have seen that energy is required to remove an electron from an atom. Removal of an additional electron or electrons from an atom requires even more energy due to the increased attraction by the nucleus for the remaining electrons. Hence it is always not possible for the atoms to form bonds by the transferrance of electrons.

Carbon has an electronic configuration 2, 4 *i.e.*, two electrons in the first energy level and four electrons in the second energy level.

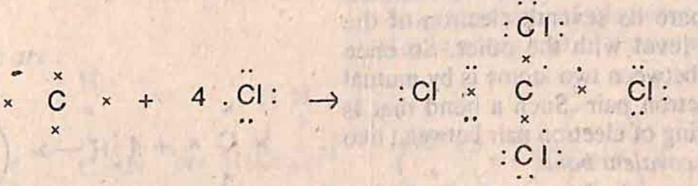
In the compound carbon tetrachloride each carbon atom is attached to four chlorine atoms. In case there is an electrovalent bond in carbon tetrachloride then four outer electrons of carbon would be transferred to each of the four chlorine atoms. But the energy required to remove four electrons from the carbon atom is very high so it does not occur. On the other hand carbon atom attains the configuration of nearest noble gas neon, by sharing the electron with chlorine atoms. This sharing of electrons between two atoms require much less energy than the transferrance of electrons. So carbon tetrachloride molecule is formed by the sharing of electrons between an atom of carbon and chlorine atom. This *each shared pair of electron between two atoms is known as covalent* bond. Hence C Cl_4 consists of four covalent bonds.

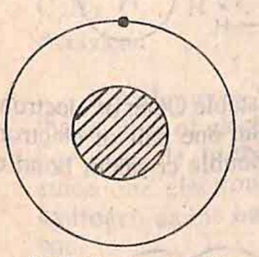
Because of this sharing of electrons, each atom involved in the formation of carbon tetra-chloride molecule *i.e.*, an atom of carbon and four chlorine atoms acquire a stable Octet of electrons.

Let us consider a few examples of homoatomic molecules. ***For example :*** Hydrogen molecule consists of two atoms of hydrogen atoms.

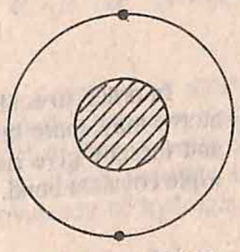
Each hydrogen atom has one electron orbitting around its nucleus. So to attain the configuration of the nearest noble gas, heluim atom it needs to have one more electron. With this electron it attains the configuration of helium atom. This it gets by sharing an electron from another atom. Two hydrogen atoms form *covalent molecules* by sharing their electrons with each other.

[Consider the formation of a chlorine molecule consisting of two atoms of chlorine. Now chlorine has the electronic configuration of 2, 8, 7. So each atom is one electron short of the nearest noble gas, argon. Electron transfer is not possible because the atoms are not complementary but are identical. Hence once again they complete the Octet of outer electrons by sharing one electron.

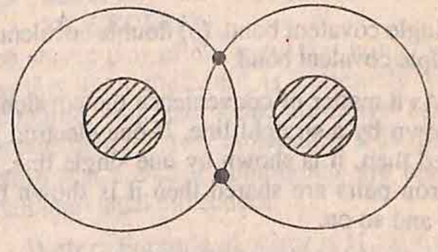




HYDROGEN ATOM

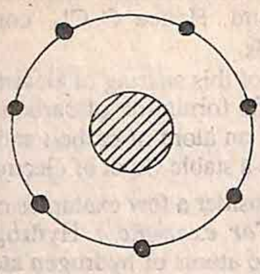


HELIUM ATOM

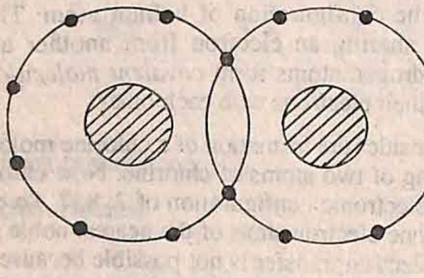


HYDROGEN MOLECULE,

TWO HYDROGEN ATOMS
SHARE ELECTRONS



Chlorine atom



Chlorine molecule

A covalent bond is formed when atoms share a pair of electrons so that both the atoms acquire stable Octet in the outermost energy level.

So when atoms of chlorine come close to each other, each one share its seventh electron of the outermost energy level with the other. So once again the bonding between two atoms is by mutual sharing of the electron pair. Such a bond that is formed by the sharing of electron pair between two atoms is known as *covalent bond.*

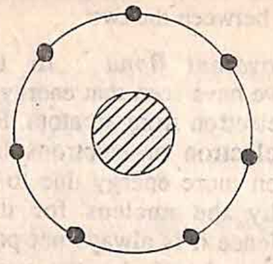
Depending upon the number of electron pairs being shared between two atoms, the bond may be

(*a*) single covalent bond, (*b*) double covalent bond, (*c*) triple covalent bond.

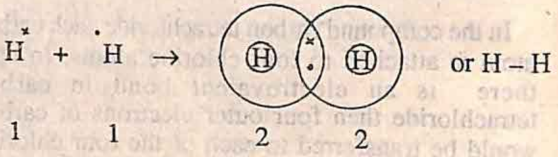
As a matter of convenience the covalent bond is shown by a straight line. If one electron pair is shared then, it is shown by one single line, if two electron pairs are shared then it is shown by two lines and so on.

4.4 SINGLE COVALENT BOND

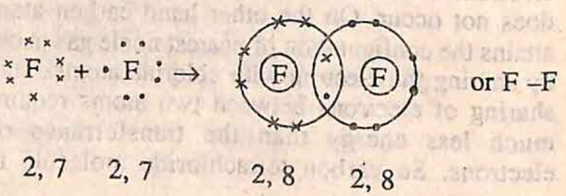
Hydrogen has one electron each and has to achieve a stable duet of electron so as to have helium configuration by sharing one electron.

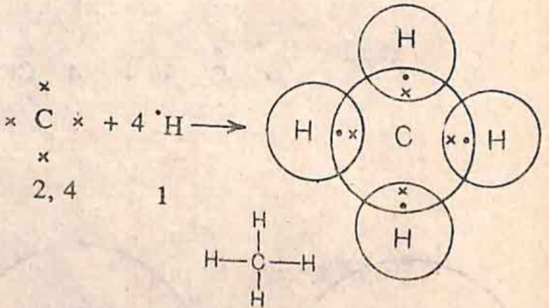


Chlorine atom

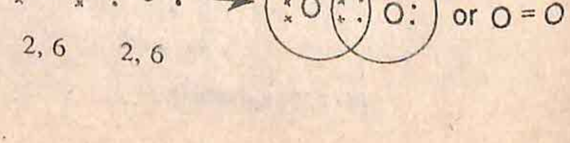


Electronic Configuration



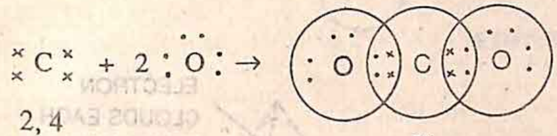


In order to achieve a stable Octet of electrons, atoms may share more than one pair of electrons and thereby give rise to double covalent bond or triple covalent bond.



Double Covalent Bond : Oxygen has six valence electrons and in order to attain the stable electronic configuration of nearest noble gas neon, it shares two electrons with other oxygen atom and thereby form a double covalent bond.

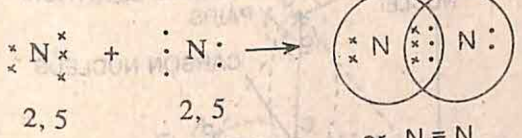
In the formation of carbon dioxide molecule, carbon has four valence electrons and oxygen has six valence electrons. So each carbon atoms share two electrons with each of the oxygen atoms and form two *double covalent bonds*.



or

$O = C = O$

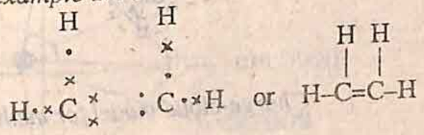
Triple Covalent Bond : Consider the formation of nitrogen molecule. Nitrogen atom has five valence electrons so requires three electrons to attain the stable Octet of neon gas. As a result the two nitrogen atoms share three electrons each and form a triple covalent bond.



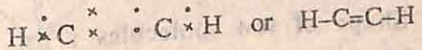
or $N \equiv N$

A few more example are :

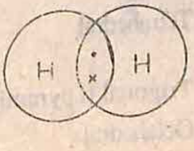
C_2H_4
Ethylene



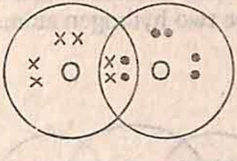
C_2H_2
Acetylene



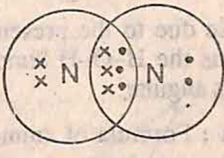
Covalency : The covalency of an atom is the number of electron pairs that the atom shares. *For example.* In the case of *hydrogen molecule*, since one electron pair is shared between two hydrogen atoms hence covalency of hydrogen is one.



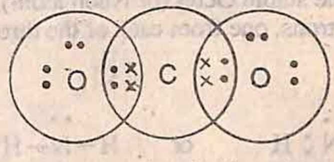
In the case of oxygen molecule two electron pairs are being shared between two oxygen atoms hence covalency of oxygen is two.



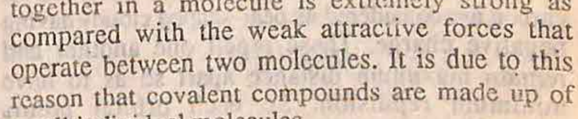
In the case of nitrogen molecule three electron pairs are shared between two atoms hence covalency of nitrogen is three.



In the case of carbon dioxide molecule covalency of carbon is four and that of each oxygen atom is two.



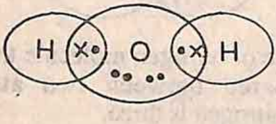
The electron pair that is shared between two atoms are known as **bond pair** of electrons and the electron pair of the valence energy level that are not shared by an atom are known as **lone pair** of electrons. So in carbondioxide molecule carbon has only lone pair of electrons but each oxyen atom has two lone pair of electrons.



4.5 SHAPES OF SOME COVALENT MOLECULES

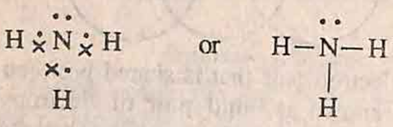
The shared pair of electrons that holds two atoms together in a molecule is extremely strong as compared with the weak attractive forces that operate between two molecules. It is due to this reason that covalent compounds are made up of small individual molecules.

Water: Formula of water is H_2O. An oxygen atom contains six electrons in its outer energy level. Two orbitals contain only one electron and are thus available for bonding [orbital is the space in an atom around the nuclei where the probability of finding electron is maximum]. The other two orbitals contain two electrons each. Such a pair of electrons are known as lone pair. So the orbitals containing an unpaired electron share two electrons from each of the two hydrogen atoms and stable its Octet.



In this case due to the presence of two lone pair of electrons the H–O–H bond angle is 105°. The molecule is angular.

Ammonia : Formula of ammonia is NH_3. In this case the element nitrogen is covalently bonded to three hydrogen atoms. Nitrogen atom has five electrons in the outermost energy level and can achieve the stable Octet (of Neon atom) by sharing three electrons, one from each of the three hydrogen atoms.



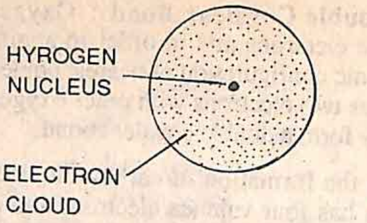
$$\mathrm{H}\,\overset{\bullet}{\times}\,\underset{\underset{\mathrm{H}}{\times\bullet}}{\overset{\bullet\bullet}{\mathrm{N}}}\,\overset{\bullet}{\times}\,\mathrm{H} \qquad \text{or} \qquad \mathrm{H}-\underset{\underset{\mathrm{H}}{|}}{\overset{\bullet\bullet}{\mathrm{N}}}-\mathrm{H}$$

Three of the orbitals contain only one electron and so are available for bonding with three spherically occupied hydrogen orbitals. The fourth orbital of nitrogen is filled as it contains two electrons. So ammonia is *tetrahedral* molecule with a lone pair of electrons being occupied by one of the four orbitals.

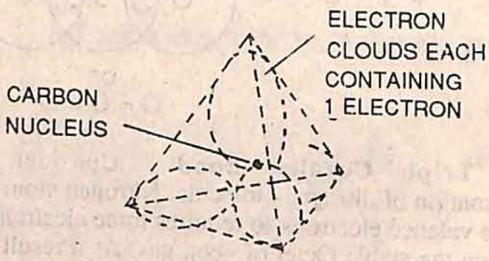
[Due to the presence of this lone pair electron the shape is not tetrahedral and pyramidal] Hence H–N–H bond angle in ammonia is 107°.28′

Methane : The four valence electrons of carbon are considered as pear-shaped clouds having negative charge. These repell one another and remain maximum distance apart so as to have minimum repulsion. So each orbital points towards the corners of a regular tetrahedron as shown in Fig.

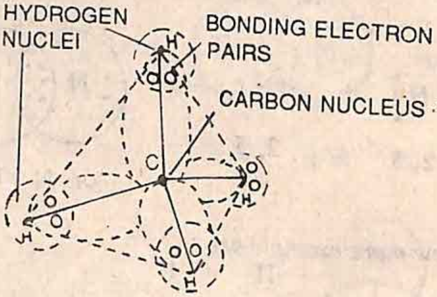
During the formation of methane each spherical orbital of four hydrogen atoms overlap with the pear shaped orbital of carbon and thereby four covalent bonds are formed and the shape of the molecule becomes tetrahedral.



Orbital cloud for the hydrogen electron



Orbital clouds for the four outer electrons of carbon

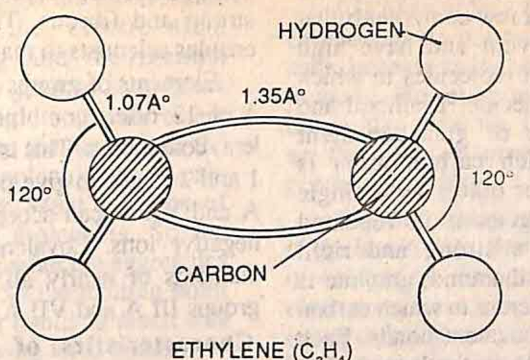


Charge-cloud model for methane

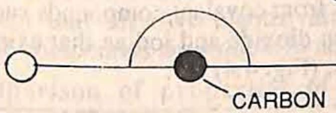
Shape of few molecules :

Formula	*Shape*	*Bond angle*
C_2H_2	Linear	180°
C_2H_4	Planar triangular	120°
$[Cu(NH_3)_4]^{++}$	Square Planar	90°
CH_4 CCl_4 $SiCl_4$	Tetrahedral	109°28′
PCl_5	Trigonal bipyramidal	90°, 120°
SF_6	Octahedral	90°

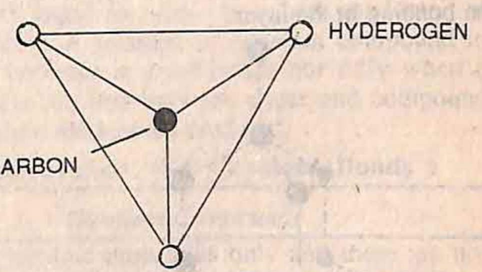
Shape of Some Simple Molecules/ions :

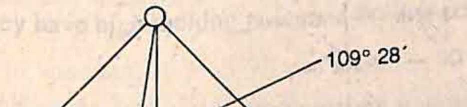


ETHYLENE (C_2H_4)



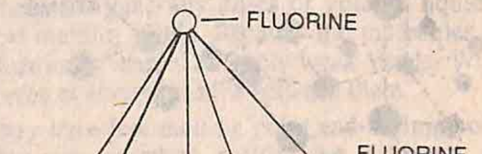
LINEAR (CO_2)





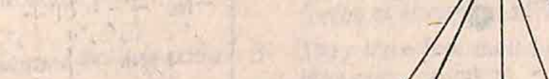
CCl_4

TETRAHEDRAL



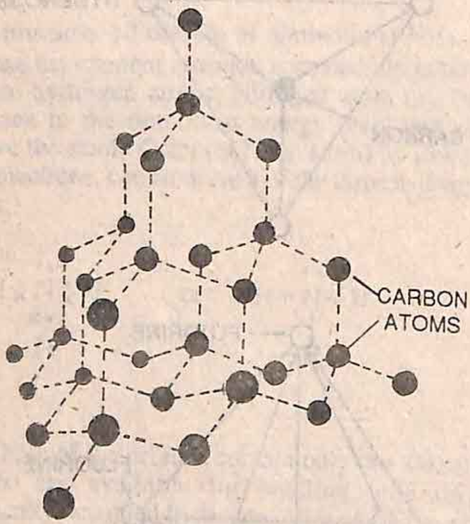
SF6

OCTAHEDRAL



SQUARE PLANAR $[Cu(NH_3)_4]^{++}$

Giant Molecules : *Diamond and Graphite* : Unlike the simple covalent compounds that we have studied so far there are a few compounds that do not dissolve in any solvent and have high melting point. They are giant molecules in which the whole crystal is one molecule. Diamond and graphite are the examples of giant covalent molecules. In diamond each carbon atom is tetrahedrally attached to four others by a single covalent bond. The arrangement is repeated throughout so as to give a strong and rigid structure to it. In contrast to diamond, graphite is soft because it has a layer structure in which carbon atoms are joined by strong covalent bonds. Each carbon is joined to three other in the largers. The bonding between the layers is much weaker than the bonding in the layer.



THE STRUCTURE OF DIAMOND

Polymers : Carbon atom has a property to combine with atoms of its own elements to form long chains. As a result of it long thread like molecules having high molecular weight are formed. Such molecules are made up of small repeating units,

$$-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-$$

Polythene or Polyethylene

[Thousands of $-CH_2-CH_2-$ Units strung together by Covalent C – C bonds]

called *monomers*, and are known as *polymers.*

A few examples are :

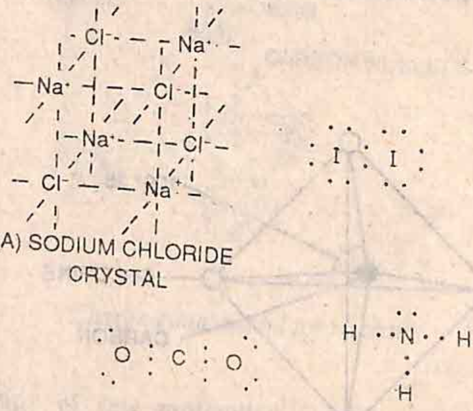
Polyethylene made up of *ethylene molecules* *Polyvinyl Chloride* (PVC) made up of Vinyl Chloride molecules other examples of polymers are nylon, Bakelite, Polyester etc. Polymers are tough, strong and fibrous. This structure and property enables scientists to make sheets and fabrics.

Elements of groups 1, 2 and VI A of 16 and VII A of 17 often combine to form ionic (electrovalent) compounds. This is because the atoms of groups 1 and 2 ionise to give positive ions and those of VI A and VII A can accept electrons readily to form negative ions. Covalent Compounds are made of elements of nearly all groups but particularly of groups III A and VII A of the Periodic Table

Characteristics of compounds containing Electrovalent and Covalent Bonds :

Electrovalent compounds such as sodium chloride, magnesium chloride with their huge rigid crystal lattice might be expected to have markedly different properties from covalent compounds such as ammonia, carbon dioxide and iodine that exists as small molecules. (Fig. 4.4)

Contrast between an ionic lattice and individual molecules (*a*) lattice structure for sodium chloride, (*b*) an iodine molecule, (*c*) a carbodioxide molecule, (*d*) an ammonia molecule.



(*a*) *Physical State* : Since there are electrostatic forces of attraction in the case of electrovalent compounds so they are solids at room temperature but covalent compounds are generally gases (CO_2, NH_3) or liquid or louring melting solids.

(*b*) *Melting Point* : Electrovalent compounds like sodium chloride are solids with high melting point. Covalent compounds are generally gases (scuh as Carbon dioxide, ammonia), liquids (such as Chloroform, Carbon tetrachloride) or low melting solids (such as iodine). This is due to the fact that the electrostatic forces between ions in the electrovalent compounds are very strong and the thermal agitation of the ions should be great so as to break down the crystal lattice. So high temperatures are required to melt the crystal. On the other hand the forces holding the covalent molecules are weak. There are only weak vander Waal's forces of attractions between the two molecules. So the thermal agitation needed to separate the aggregates of covalent molecules is low. So the melting point of compound having covalent bonds is much less as compared to compounds having ionic bonds.

(c) *Solubility* : Electrovalent compounds are soluble only in water. This is because water is capable of breaking up the crystal lattice into ions. Covalent compounds are insoluble in water *i.e.*, solvent that dissolve electrovalent compounds. These compounds are soluble in non-polar solvents like carbontetra chloride, petrol and paraffin.

(d) *Electrical Properties* : Electrovalent compounds do not conduct electricity in the solid state, because in the solid state the ions are not free to move. However, when dissolved in water or in the molten state the lattice is destroyed, the ions are free to move around, and they are then able to conduct electricity.

Covalent compounds do not contain ions even in the liquid state, therefore they are bad conductor of electricity. Pure water is a bad conductor of electricity, but a solution of an electrovalent solid in water is a good conductor. The water is able to break down the crystal lattice and the ions are now free to move. A solution of covalent compound such as sugar in water is a bad conductor of electricity. A solution of covalent compound in water becomes a good conductor only when a chemical reaction between water and compound takes place and ions are produced.

Comparison of properties of Compounds having Electrovalent and Covalent Bonds :

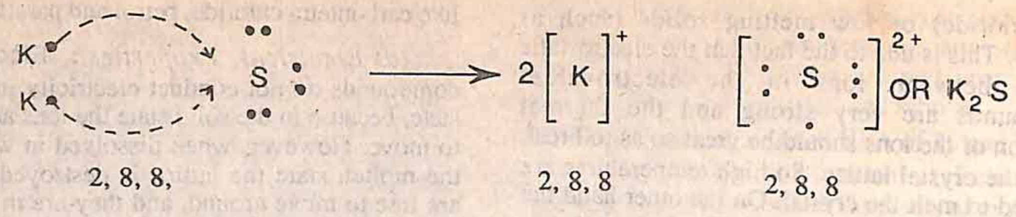
Electrovalent Compounds	*Covalent Compounds*
1. They contain oppositely charged ions only and there are no molecules in them. e.g., Sodium Chloride, Magnesium Oxide	1. They contain molecules only and there are no ions present in them *e.g.*, iodine ; carbondioxide; ammonia.
2. There are solids at room temperature. The crystalline solids consists of ions that are held together by strong electrostatic forces of attraction.	2. These are generally gases or volatite liquid or low melting solids. Because the molecules are electrically neutral and only weak Vander Waals forces of attraction exist between them.
3. They have high melting point and boiling point.	3. They have low melting point and boiling point. However covalent solids like diamond and graphite are exceptions.
4. They conduct electricity in the molten state and in aqueous solution and are therefore electrolytes. This is due to the presence of free ions in the molten state and in aqueous solution.	4. They do not conduct electricity in the molten state or in aqueous solution. Hence they are non electrolytes.
5. They are soluble in water but insoluble in organic solvents such as petrol, paraffin, ether, etc.	5. They are mostly soluble in organic solvent such as ether, paraffin and petrol but sparingly soluble in water
6. They are non directional.	6. They are directional.

Experiment 4.1 Potassium sulphide is a typical alkali metal sulphide. It is put into water and the resulting solution conducts an electric current. Comment on the nature of compound.

Potassium has atomic number 19. So its electronic configuration is 2, 8, 8, 1. It can attain the configuration of nearest noble gas argon by donating an electron.

Sulphur has atomic number 16. So its electronic configuration is 2, 8, 6. It can attain the electronic configuration of the nearest noble gas argon by gaining two electrons.

Hence two atoms of potassium donate one electron each to sulphur and the electrovalent compound potassium sulphide is formed.



Experiment 4.2. *Electrovalent bonds confer electrical conductivity to substances.* Take a solution of sodium chloride in water and dip the two electrodes in the salt solution. Connect these two electrodes to a battery via electric bulb as shown in the figure. On pressing the key the bulb glows. This phenomenon confirms that the solution of sodium chloride furnishes ions that are free to move and conduct electricity. Other examples of such compounds are $MgCl_2$; K Br, $CuSO_4$.

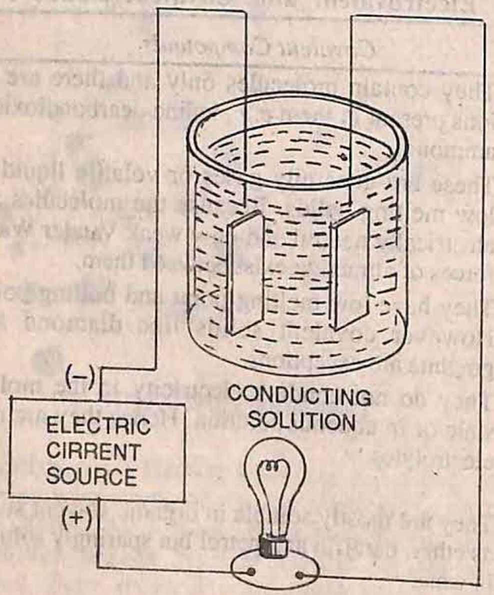
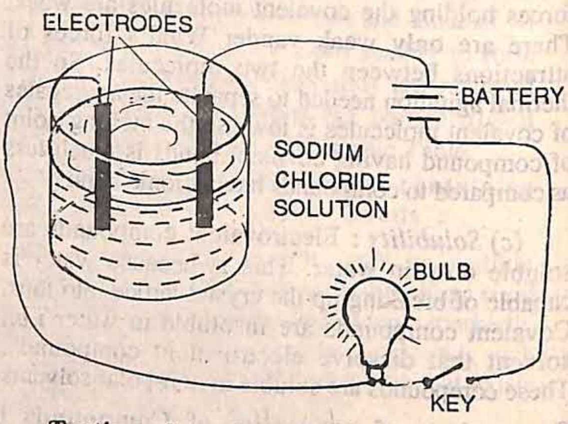


Fig 4.5 : Aqueous solution of Electrovalent compounds conduct electricity. Conduction through the solution is signalled by the glowing bulb in the path.

Experiment 4.3 Distinguish between a electrovalent and covalent compound.



Testing for electrical conductivity of solutions. Take a beaker half filled with distilled water. Insert two electrodes into the beaker and set up the apparatus as shown above in the figure. Press the key. What happens ? The bulb does not glow.

What do you conclude about the electrical conductivity of the distilled water ?

It is concluded that distilled water is not a good conductor of electricity.

Now add a pinch of sodium chloride in water. Stirr the solution and again press the key. Now the bulb glows. What do you conclude about the electrical conductivity of sodium chloride solution ?

It is concluded that an aqueous solution of sodium chloride conducts electricity *i.e.*, sodium chloride aqueous is a good conductor of electricity.

Now replace the sodium chloride solution by carbon tetrachloride. Press the key. What do you observe ? The bulb does not glow this time. From this it can be concluded that carbon tetrachloride is not a conductor of electricity.

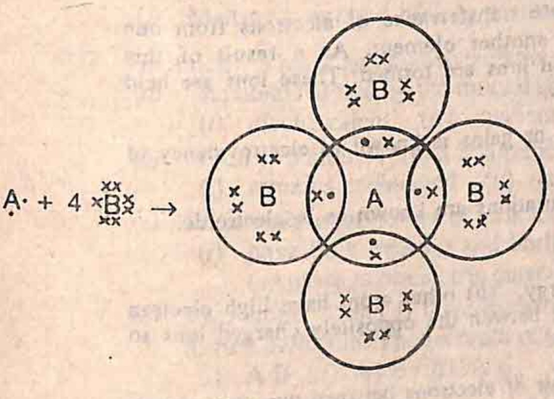
It means compounds like NaCl are good conductors of electricity. *i.e.*, electrovalent compounds conduct electricity whereas those like carbon tetrachloride *i.e.*, covalent compounds do not conduct electricity.

Experiment 4.4. Solubility of sodium chloride in benzene and water at room temperature.

Take about 5 *ml* of water in one test tube and 5 *ml* of benzene in another test tube. Add a pinch of sodium chloride to each test tube. Shake for about two minutes and observe. It is observed that sodium chloride dissolves in the test tube containing water but does not dissolve in the test tube containing benzene. Hence we can conclude that electrovalent compounds are soluble in water.

Problem : An element A has four valence electrons in its atom and it combines with another element B having seven valence electrons in its atom. What is the nature of the chemical bond in the compound so formed ? Give the electronic structure of its molecule. Write down its formula.

Solution : Since atom `A' has four valence electrons it requires four more electrons to attain the configuration of nearest noble gas. This it can neither do by gaining four electrons nor losing four electrons because a lot of energy is required for this purpose. Atom 'B' has seven valence electrons so it requires only one electron to attain the configuration of nearest noble gas. Hence four atoms of B share one electron each with an atom of A and form covalent bond.



Hence formula of the compound is AB and it is a *covalent compound.*

Problem: An element 'A' has two electrons in the outermost energy level of its atom and it combines with an element B having seven electrons in the outermost energy level of its atom. The compound so formed is soluble in water. Giving reason predict the nature of the chemical bond in the compound. State two proerties of such compounds and write down its formula.

Solution : Atoms of 'A' has two valence electrons hence in order to attain the configuration of nearest noble gas it must lose two electrons. On the other hand atoms of B has seven valence electrons so to attain the configuration of nearest noble gas (*i.e.*, Octet) it must gain an electron. As a result an atom of A donate one electron each to an atom of B and form electrovalent bond.

$$\overset{\times}{\text{A}}{\times} + 2\ \cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{B}}}: \rightarrow [\text{A}]^{2+}\ 2[\overset{\times}{\underset{\times}{\cdot}}\overset{\cdot}{\underset{\cdot\cdot}{\text{B}}}:]$$

or AB_2

Since the bond is electrovalent hence the AB_2 compound is soluble in water.

Characteristics of AB_2 :

1. It is solid at room temp.
2. It conducts electricity in the molten state and in aqueous solution.
3. It has high melting point.

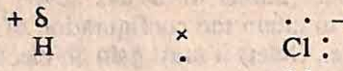
In certain compounds the same molecule can have both ionic and covalent bonds in it. ***For example*** : In the case of sodium hydroxide. There is an ionic bond between sodium and hydroxide. But a covalent bond exists between oxygen and hydrogen. $Na^+\ (O{-}H)^-$. In the same way in the case of hydrogen cyanide, an ionic bond exists between hydrogen and cyanide. But a triple covalent bond exists between carbon and nitrogen.

$$(H^+)\ [C^- \equiv N]^-$$

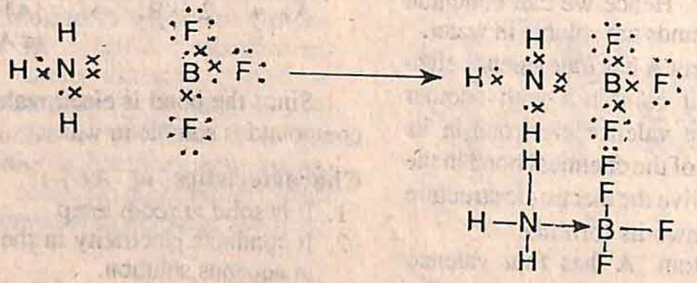
A few more examples of these types of compounds are :

Sodium Carbonate	Na_2CO_3
Calcium Carbonate	$CaCO_3$
Sulphuric Acid	H_2SO_4
Ammoniam Nitrate	NH_4NO_3
Ammoniam Chloride	NH_4Cl
Ammonium Sulphate	$(NH_4)_2SO_4$

Polar Covalent : We know a covalent bond is formed by the sharing of electron pair between two atoms. a bond is hundred percent covalent bond if the electron pairs are being shared between atoms of the same element. ***For example*** : Cl_2 ; H_2 ; N_2 ; O_2. Because in such case the shared electron pair lies exactly in between the two. But if the electron pair is being shared between atoms of different elements then the shared electron pair lies slightly more towards one atom, which is more electro-negative [Electro-negativity is the power with which a bonded atom in a molecule attracts the shared pair of electrons towards itself]. As a result that atom towards which the shared electron pair lies more gets a slight negative charge and the other acquires a slight positive charge. Such a bond is known as *polar covalent bond*. ***For example:***



Coordinate covalent bond : We know during covalent bond formation an electron pair is shared between two atoms and each atom contributes towards sharing *i.e.*, both the atoms provide one electron each for sharing. But in certain cases the shared pair of electron is provided by one atom. The atom that provides the lone pair of electron for sharing is known as *donor* and the other that simply takes part in sharing is called *acceptor*. The bond is shown by an arrow (→) pointing from the donor to the acceptor. ***For example:***,



SUMMARY

4.1 The force that holds atoms together in a molecule is known as Bond. When the two atoms approach near each other there is a decrease in energy and the energy that is released is known as Bond energy.

4.2 This decrease in energy is responsible for the formation of compound.

4.3 In many cases when elements form compounds their atoms either lose, gain or share electrons in order to have stable Octet just similar to next higher or lower noble gas atom.

4.4 An electrovalent bond involves the complete transferrance of electrons from one atom of an element to another atom of another element. As a result of this transferrance of electrons, oppositely charged ions are formed. These ions are held together by electrostatic forces of attractions.

4.5 The number of electrons that an atom loses or gains is known as electrovalency of that atom.

4.6 Species that have the same electronic configurations are known as Isoelectronic.

4.7 In general electrovalent bond is formed when

(*a*) one of the atoms have low ionistion energy. (*b*) other atom have high electron affinity. (*c*) strong forces of attraction exist betwen the oppositely charged ions so formed.

4.8 A covalent bond involves the sharing of a pair of electrons between two atoms, each atom contributes one electron to the shared pair.

4.9 Depending upon the number of electron pairs being shared between two atoms, the bond may be

(*a*) single covalent bond, (*b*) double covalent bond, (*c*) triple covalent bond.

4.10 Covalency of an atom of an element is the number of electron pairs that the atom shares.

4.11 The lone pair of electron that is shared between two atoms are known as bond pair of electrons and the electron pair of the valence energy level that are not shared by an atom are known as lone pair of electrons.

4.12 Polymers are the compounds having high molecular weight. These are griant molecules and are formed by the combination of small repeating units called monomers.

4.13 Electrovalent compounds are generally solid at room temperature, soluble in water ; have high melting point and conduct electricity in the molten state and in aqueous solution.

4.14 Covalent compounds are generally gases at room temperature ; or low melting solids or volatile liquids. They do not conduct electricity in molten or solid state and are insoluble in water.

4.15 The properties of a substance depend upon the arrangement of the atoms in the molecules of the substance.

QUESTIONS

4.1 Complete the following statement :

(*a*) .... is a force that hold atoms together in a molecule.

(*b*) The two types of bonds are .... and ....

(*c*) An electrovalent bond is formed by the .... of electrons.

(*d*) During electrovalent bond formations one of the atoms should have .... ionisation, energy and other should have .... electrons affinity.

(*e*) Electrovalent compounds are soluble in .... solvents.

(*f*) Covalent bond is formed by .... of the electrons.

(*g*) Covalent compounds are soluble in......................

(*h*) During the formation of an electrovalent compound, the metallic element .... electrons while non–metallic element .... electrons.

(*i*) Graphite which exists in layers is an example of .... compounds.

(*j*) Melting point and boiling point of electrovalent compound is ....

4.2 Put a tick (√) mark against the suitable choice :

(*a*) The bond formed by the mutual sharing of electrons is :
(*i*) electrovalent. (*ii*) covalent. (*iii*) metallic.

(*b*) When two atoms form a bond, the energy of the system :
(*i*) remains unchanged. (*ii*) increases. (*iii*) decreases. (*iv*) may increase or decrease.

(*c*) Which of the properties of ionic compounds given below is not correct ?
(*i*) have high melting and boiling point. (*ii*) soluble in water. (*iii*) ionic solids permit the passage of electric current. (*iv*) melt permits the passage of electric current.

(*d*) Two atoms A and B have electronic configurations 2, 8, 8, 2 and 2, 8, 7 respectively and form a molecule. The formula of the molecule will be :
(*i*) AB (*ii*) A_2B (*iii*) AB_2 (*iv*) BA_2.

(*e*) A cation is formed by
(*i*) sharing a pair of electrons. (*ii*) accepting an electron. (*iii*) losing an electron. (*iv*) gaining a proton.

(*f*) Electrovalent compounds conduct electricity when dissolved in water because :
(*i*) the ions are free in solution (*ii*) crystals structure persists. (*iii*) electrons are supplied by water. (*iv*) ions aggregates in solution.

(*g*) Shape of ammonia molecule is
(*i*) linear. (*ii*) pyramidal. (*iii*) angular. (*iv*) octahedral.

(*h*) $[Cu(NH_3)_4]^{++}$ is
(*i*) square planar. (*ii*) linear. (*iii*) tetrahedral. (*iv*) octahedral.

(i) SF_6 is

(i) square planar (ii) linear (iii) tetrahedral (iv) octahedral

(j) CClu is :

(i) square planar (ii) linear (iii) tetrahedral (iv) octahedral

(k) Carbondioxide molecule is

(i) trigonal bipyramid (ii) linear (iii) planar (iv) Tetrahedral

(l) Phosphorus pentachloride is

(i) trigonal bipyramid. (ii) octahedral. (iii) Planar. (iv) square planar.

(m) Covalent bonds are

(i) directional. (ii) non-directional. (iii) polar. (iv) weak.

(n) Which of the following does not have both ionic and covalent bond ?

(i) Sodium hydroxide (ii) Hydrogen cyanide (iii) Water (iv) Sulphuric acid.

4.3 Match the descriptions in Column B with the terms/statements in Column A :

Column A	Column B
(a) Electrovalent bond	(a) Ions are held by strong electrostatic forces
(b) Covalent bond	(b) Octahedral
(c) Chloroform	(c) linear
(d) Carbondioxide	(d) Covalent Compound
(e) Silicon tetrachloride	(e) Formed by the sharing of electrons
(f) Sulphur hexafluoride	(f) Transferrance of eletrons
(g) Sodium Chloride	(g) Square planar
(h) $[Cu(NH_3)_4]^{++}$	(h) tetrahedral

4.4 Point out the wrong statements of the following :

(a) Ionic compounds are formed by the transferrance of electrons.

(b) Covalent compounds are soluble in polar solvents.

(c) Octet rule is always helpful in explaining the structures of various molecules.

(d) Hydrogen Chloride is a electrovalent compound.

(e) A covalent bond formed by the sharing of three electron pairs between two atoms is known as triple bond.

(f) Electrovalent compounds have high melting and boiling points.

(g) Electrovalent bond exists between two atoms of the same elements when combine to form a molecule.

(h) Silicon tetra fluoride is a covalent compound.

(i) Ethylene molecule is planar.

(j) Phosphorus pentachloride is a trigonal bipyramid.

(k) Covalent bonds have definite geometry about them.

(l) The H–O–H bond angle in water is 105°.

4.5 What is a chemical bond ?

4.6 What are the ways in which an atom can attain noble gas configuration ?

4.7 What is the difference between sodium atom and sodium ion ? Will their properties be the same or different ?

4.8 What type of bonding would you expect in the following sets when two atoms A and B with the following electronic configuration combine with each other ?

(a) A = 2 ; 8, 1 B = 2, 7

(b) A = 2, 8, 8, 2 B = 2, 6

(c) A = 2, 8, 7 B = 2, 8, 7.

4.9 In the formation of a compound AB, atoms of element A lost two electrons each while atoms of element B gained two electrons each. Predict the nature of bond between A and B. List two characteristics of this compound.

4.10 Electronic configuration of an atom X is 2, 8, 7. It combines with another atom Y will electronic configuration :

(*a*) What type of bond will be formed between X and Y ?
(*b*) Write the formula of the compound formation.

4.11 An element M burns in oxygen to form an electrovalent compound MO. What compounds will be formed if this element combines with Chlorine and Sulphur ?

4.12 If X and Y have the following electronic configurations, how many atoms of the element X, will combine with one atom of Y ?

X = 2, 8, 7
Y = 2, 4.

Write down the electron dot formula of the compound formed from X and Y. List its two properties.

4.13 How is an electrovalent bond formed ? Show the formation of bond in Sodium sulphide compound. [Atomic Number of sodium = 11, atomic number of Sulphur 16].

4.14 In the given compounds find out which compound will have electrovalent bond ?
(*a*) Nacl (*b*) MgO (*c*) KCl (*d*) H_2O (*e*) HCl (*f*) $SiCl_4$
(Atomic number of : Mg = 12; Na = 11 ; Cl = 17 ; O = 8; H = 1 ; Si = 14 ; K = 19)

4.15 Write down the formula of the compound formed by the combination of an element X (atomic number 20) with element Y (atomic number 17).

4.16 How many covalent bonds do the following molecules have ?
(*a*) Ammonia (*b*) Water (*c*) Methane (*d*) Sulphur hexafluoride

4.17 Justify that the formula of Magnesium oxide is MgO and that of sodium oxide is Na_2O.

4.18 What happens when an atom of Sodium is brought close to a chlorine atom ?

4.19 Atomic numbers of Calcium, Aluminium and Oxygen are 20, 13 and 8 respectively. What are the valencies of Ca, Al and Oxygen ?

4.20 Is it equal to valence electrons or 8-valence electrons ? Write down the formula of the Calcium, oxide and Aluminium oxides.

4.21 What are polymers ? Give two examples

4.22 What is the shape of the following molecules /species ?
(*a*) Ammonia (*b*) Water (*c*) Ethylene (*d*) Phosphorus pentachloride (*e*) Sulphur hexafluoride (*f*) Ethylene.

4.23 The elements A, B, C and D have atomic numbers 7, 9, 10 and 11 respectively.Write the formulae of the compounds that you would expect to form between the following pairs of elements and also indicate the type of bonding present in them.
(*a*) A and B (*b*) B and B (*c*) A and D (*d*) C and C.

4.24 Choose the ionic compounds of the following compounds :
(*a*) Cl_2 (*b*) Mg Cl_2 (*c*) AlF_3 (*d*) NH_3 (*e*) CO_2 (*f*) CH_4 (*g*) Mg O (*h*) $CaCl_2$

4.25 You are provided with a white powder that is soluble in water. How will you show that white powder is an electrovalent compound ?

4.26 Compare the characteristics of electrovalent and covalent compounds.

□□

CHAPTER 5
CHEMICAL REACTIONS

5.1 INTRODUCTION :

Physical and chemical changes: On heating, wax melts and iodine gives violet vapours. Both these substances, on cooling, turn back into the same substance. But, on the other hand, coal and wood, on burning, give off gases and the residue that is left is quite different from the original wood and coal. So, from these experiments we can conclude that heat can bring about two types of changes :

- *(i)* One in which there is a change of physical state of the substance, e.g. melting of ice, evaporation of water, melting of wax.
- *(ii)* Other in which new substances are formed. The properties of these new substances are different from those of the parent. Based on this the changes that take place daily around us can be classified into two heads:

(a) **Physical Change:** These are the changes in which no new substance is formed. There is only a change in the physical state of the substance or the colour, e.g., melting of ice, formation of clouds, melting of wax, lightening of a bulb. During this change, there is, no change in the atoms making up the molecules.

(b) **Chemical Change:** These are the changes in which new substances are formed. The properties of the substance formed are entirely different from the parent's. During a chemical change the atoms are rearranged into other molecules.

Let us study the difference between the physical and chemical changes in order to have a clear idea of what they are:

Physical Change	Chemical Change
1. There is a change in the physical state of substance only.	Physical state of the substance may or may not undergo any change.
2. No new substance is formed.	New substances are formed.
3. It can be reversed by reversing the conditions.	In can be reversed but with a lot of difficulty.
4. Heat changes accompanying the process are usually small.	Large heat changes accompany the process.
5. The substances formed have the same chemical properties. e.g., evaoration of water.	The substances formed have different chemical properties. e.g., electrolysis of water into hydrogen and oxygen.

5.2 CAUSE OF CHEMICAL CHANGE

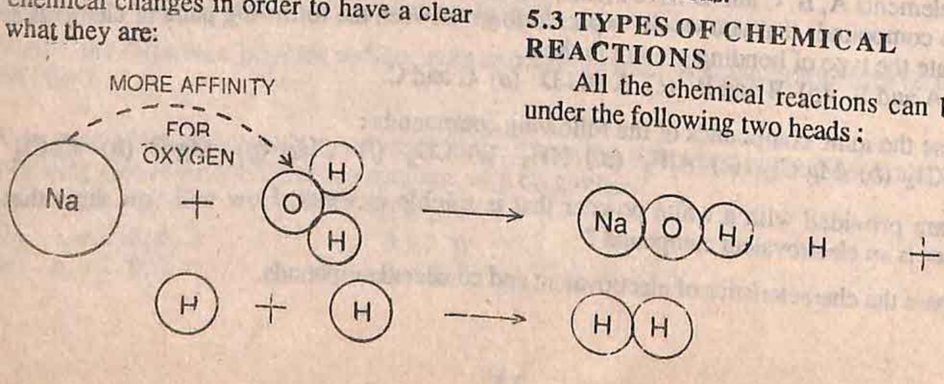
According to Newton everybody attracts every other body with a force of attraction. This force of attraction depends upon the mass and the distances between them. In the same way when two substances react with each other, there exists a chemical force of attraction between them under the given conditions of temperature, pressure and concentration. This force of attraction by virtue of which a substance enters into a chemical combination with another is called chemical affinity.

For example : Sodium reacts with water to give sodium hydroxide and hydrogen gas is given out. This is because of the greater affinity of sodium for oxygen of water as compared to that of hydrogen for oxygen.

So *chemical affinity* is the force behind all chemical reactions.

5.3 TYPES OF CHEMICAL REACTIONS

All the chemical reactions can be classified under the following two heads :



(*a*) Reactions involving a change in the same substance.

(*b*) Reactions that occur between two or more different substances.

5.3.1 Reactions involving the same substance can be of two types ?

(*a*) Rearrangement Reactions (*b*) Decomposition reactions

(*a*) **Rearrangement Reactions (Isomerism) :** These reactions involve a change in the internal structure of a substance resulting in the formation of a new substance having different structure and different properties.

For Example : Ammonium Cyanate on heating gives urea.

$NH_4\,CNO \xrightarrow{heat} NH_2\,CONH_2$

Ammonium Cyanate Urea

In these the molecular formula is the same but bonds are rearranged to give new compound. Such compounds having the same molecular formula but different structural arrangements are called **isomers.** *Isomerism* is more common in organic chemistry.

(*b*) **Decomposition Reactions :** Decom position is the process of breaking up of a substance into two or more compounds. This process of breaking up of compounds can be brought about by the effect of heat, light or electricity. Hence the decomposition reactions can be of the following types :

(*i*) **Thermal decomposition :** Thermal decomposition reactions are those that are brought about by the actions of heat.

For example :

$CaCO_3\,(s) \xrightarrow{heat} CaO(s) + CO_2(g)$

Calcium Carbonate Calciumoxide Heat.

$2\,HgO(s) \xrightarrow{heat} 2\,Hg(l) + O_2\,(g)$

Mercuric Oxide Mercury Oxygen

$NH_4\,Cl\,(s) \xrightarrow{heat} NH_3(g) + HCl(g)$

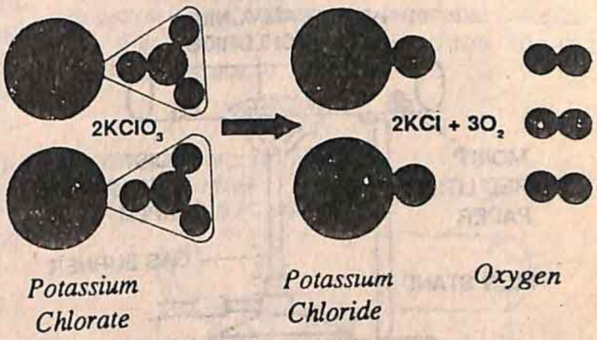
Ammonium chloride Ammonia hychogen chloride

$Mg(OH)s\,(s) \xrightarrow{heat} MgO(s) + H_2O(g)$

Magnesium hydroxide Magnesium oxide Water

$2\,K\,Cl\,O_3\,(s) \xrightarrow{heat} 2K\,Cl\,(s) + 3O_2\,(g)$

THERMAL DECOMPOSITION REACTION



(*ii*) **Electrical decomposition or Electrolysis:**

In case the decomposition is carried out by passing electricity then the reaction is known as electrical decomposition :

$2\,Na\,Cl\,(l) \xrightarrow{Electrolysis} 2Na^{(l)} + Cl_2\,(g)$

Sodium Chloride Sodium Chlorine

$2H_2O\,(l) \xrightarrow{Electrolysis} 2\,H_2(g) + O_2(g)$

Water Hydrogen Oxygen

$Ca\,Cl_2(l) \xrightarrow{Electrolysis} Ca(s) + Cl_2(g)$

Calcium Chloride Calcium Chlorine

$Mg\,Cl_2(l) \xrightarrow{Electrolysis} Mg(l) + Cl_2(g)$

Magnesium chloride Magnesium Chlorine

(*iii*) **Photochemical decomposition:** Photochemical decomposition reactions are those that are carried out by the effect of light. In photography the film is exposed by focussing an image upon the light sensitive photographic emulsion (Ag Br) on the film. The silver bromide on receiving the light is reduced by light.

$Ag\,Br \longrightarrow Ag^+\,Br-$

Silver Bromide Silver Bromide ion

$Br^- + photon \longrightarrow Br + e^-$

(light) (Absorbed on gelatin)

$Ag^+ + e \longrightarrow Ag.$

Experiment 5.1 : To demonstrate thermal decomposition.

(*a*) Take some solid ammonium chloride in a hard glass tube open at both end. Ammonium chloride is placed in the middle of the tube as shown in Fig. 5.1. At one end of the tube place moist red litmus paper and on the other end of tube place moist blue litmus paper.

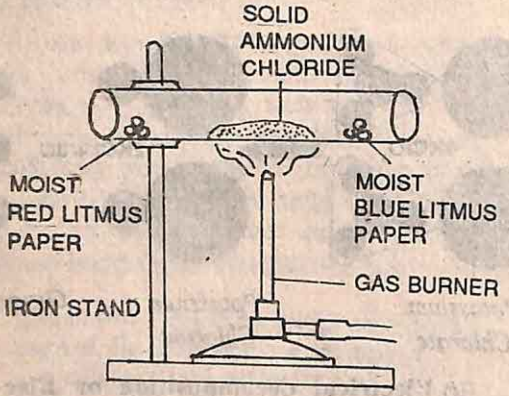


Fig. 5.1 Decomposition of Ammonium Chloride

Now start heating ammonium chloride. On heating it decomposes into ammonia gas and hydrogen chloride gas. And after sometime blue litmus paper becomes red indicating the presence of hydrogen chloride gas and red litmus paper becomes blue indicating the presence of ammonia gas.

Experiment 5.2 Take a little of zinc carbonate in a hard glass boiling tube. Close the mouth of the tube with a rubber cork fitted with a delivery tube bent at an angle. Dip the other end of the delivery tube in lime water as shown in Fig. 5.2. Heat the hard glass boiling tube. After sometime Zinc carbonate decomposes. Evolution of bubbles indicate the decomposition of Zinc carbonate. After sometime lime water turns milky. Do you know why does lime water turn milky ?

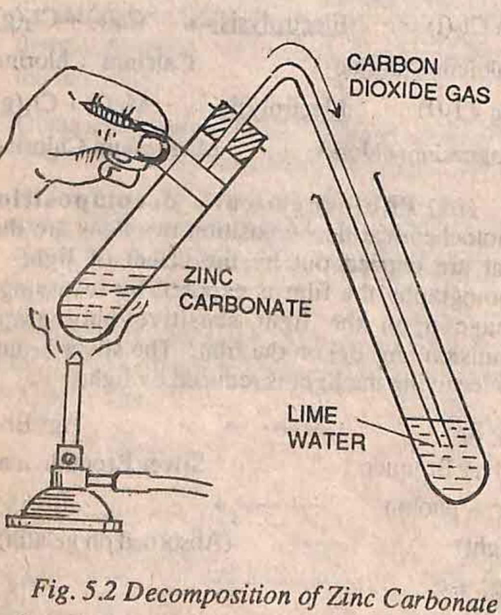


Fig. 5.2 Decomposition of Zinc Carbonate

This is because carbon dioxide reacts with lime water, and calcium carbonate is obtained. Calcium carbonate is insoluble in water. Hence milkiness appears.

If we continue passing carbon dioxide for a sometime, the milkiness disappears. This is due to the formation of calcium bicarbonate which is soluble in water.

Experiment 5.3 : Decomposition of water by electric current.

Take Hoffmann's voltameter. Fill the two side tubes completely with water containing a little sulphuric acid. Pure water does not conduct electric current and a little acid is added to enable it to do so. The acid does not change or get used up during the process of electrolysis. Close the tap of the side tubes and from the bottom insert two electrodes (platinum or iron or any metal). Platinum is preferred because it does not react during electrolysis. Connect the

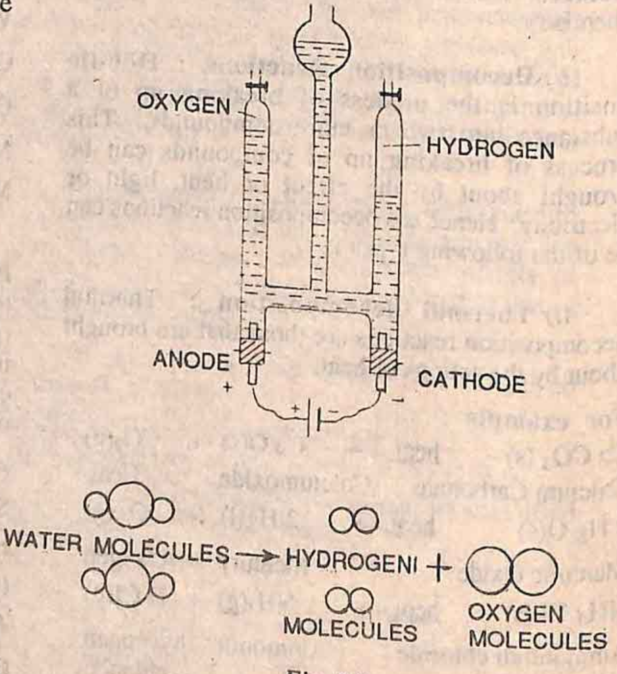


Fig. 5.3

Decomposition of Water by Electric Current electrodes to a battery. On passing electric current water splits up into two gases. Twice as much hydrogen gas collects above the cathode as oxygen above the anode.

Reactions involving two or more different substances are of the following types :

(*a*) **Combination Reactions** : Combination reactions are those in which two or more elements or compounds combine to form a single compound. For example, Magnesium burns in the presence of oxygen to give magnesium oxide.

$2\,Mg(s) \quad + \quad O_2(g) \rightarrow \quad 2\,MgO(s)$
Magnesium Oxygen Magnesium oxide.

Here two elements combine to form a compound. Examples of this type are :

$H_2(g) \quad + \quad Cl_2(g) \rightarrow \quad 2\,HCl(g)$
Hydrogen Chlorine Hydrogen chloride.

$2\,Hg(l) \quad + \quad I_2(s) \rightarrow \quad Hg_2I_2(s)$
Mercury Iodine Mercurous iodide

$C(s) \quad + \quad O_2(g) \rightarrow \quad CO_2(g)$
Carbon Oxygen Carbon dioxide gas

$4\,Al(s) \quad + \quad 3\,O_2(g) \rightarrow \quad 2\,Al_2O_3(s)$
Aluminium Oxygen Aluminium oxide.

Experiment 5.4 : Weigh 0·5 gram of mercury and 1·0 gram of iodine. Mix the two in mortar and pestle. Grind the two thoroughly. After sometimes a green substance mercurous iodide is obtained.

Experiment 5.5 : Hold a piece of magnesium wire between a pair of tongs over a flame. After sometime it burns with a dazzling light. A white substance Magnesium oxide is obtained.

There are a number of reactions in which a compound and an element combine together to form a new compound. Examples of this type are :

$2SO_2(g) \quad + \quad O_2(g) \rightarrow \quad 2\,SO_3(g)$
Sulphur dioxide Oxygen Sulphur trioxide

$C_2H_4(g) \quad + \quad H_2(g) \rightarrow \quad C_2H_6(g)$
Ethene Hydrogen Ethane

$2FeCl_2(aq) \quad + \quad Cl_2(g) \rightarrow \quad 2\,FeCl_3(aq)$

There are a number of reactions in which two or more compounds combine to give a single compound. Examples of this type are :

$NH_3(g) \quad + \quad HCl(g) \rightarrow \quad NH_4Cl(s)$
Ammonia Hydrogen chloride Ammonium chloride

$CaO(s) \quad + \quad H_2O(l) \rightarrow \quad Ca(OH)_2(s)$
Calcium oxide Water Calcium hydroxide

$Na_2O(s) \quad + \quad H_2O(l) \rightarrow \quad 2\,NaOH(aq)$
Sodium oxide Water Sodium hydroxide

Experiment 5.6 : Take a glass tube open at both ends. Hold it horizontally in the iron stands as shown in the Fig. 5.4. At one end place a cotton dipped in ammonia (aqueous) solution and at the other end place a cotton dipped in hydrochloric acid.

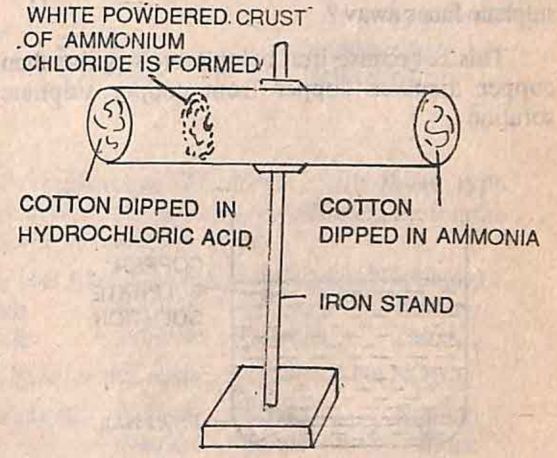

WHITE POWDERED CRUST OF AMMONIUM CHLORIDE IS FORMED
COTTON DIPPED IN HYDROCHLORIC ACID
COTTON DIPPED IN AMMONIA
IRON STAND

Fig. 5.4

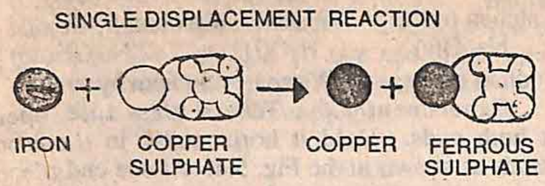
After sometime white powdered ammonium chloride is formed near the cotton dipped in hydrochloric acid. Do you know why the powder (Ammonium chloride) is formed near Hydrochloric acid end and not near the ammonia end ? This is because ammonia is a lighter gas so it diffuses more readily.

Polymerisation : Polymerisation is a combination reaction in which many small molecules (of the same or different compounds) combine together to form a single large molecule. **For example** : Acetylene gas passed through a red hot tube under pressure gives benezene.

$$3\ \begin{matrix}CH\\|||\\CH\end{matrix} \xrightarrow[\text{Copper tube}]{\text{Red hot}} C_6H_6$$

Acetylene Benzene

Single displacement or Substitution reactions : In these type of reactions an element reacts with a compound and displaces one of the elements in the compound. **For example** : Take



copper sulphate solution in a beaker. Dip a needle in it. After sometime blue colour of copper sulphate fades away why the blue colour of copper sulphate fades away ?

This is because iron being more reactive than copper displace copper from copper sulphate solution

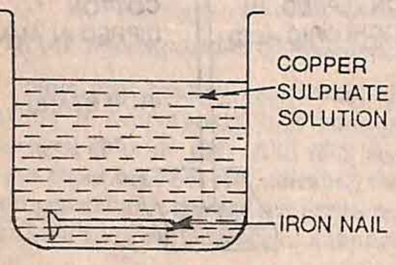


Fig. 5.5

$CuSO_4(aq) + Fe(s) \rightarrow FeSO_4(aq) + Cu(s)$

Copper sulphate   iron   Iron sulphate   Copper

$2AgNO_3(aq) + Cu(s) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

Silvernitrate   Copper   Copper nitrate   Silver

$FeSO_4(aq) + Zn(S) \rightarrow ZnSO_4(aq) + Fe(s)$

Ferrous Sulphate   Zinc   Zinc sulphate   Iron

The element that displaces another element from its compound is more reactive than that element. The activity of the metallic elements increases down the group in the periodic table. **For example** :

In the case of alkali metals

$Na < K < Rb < Cs$ Increases down the group.

But in non-metals the reactivity is in the order.

$F_2 > Cl_2 > Br_2 > I_2$ decreases down the group.

Experiment 5.7 : Set up the apparatus as shown in Fig. 5.6. Take a hard glass tube open at both ends. At one end place potassium iodide. Then place the glass wool. After this place potassium bromide. Then again place glass wool. Fix a cork at the end of the tube. Pass a delivery tube through it. Dip the other end of the tube in a hard glass boiling tube containing manganese dioxide. Add HCl Conc. to the test tube through thistle funnel. Chlorine gas is liberated immediately. As it approaches potassium bromide, reddish brown vapours of bromine are seen. After sometime violet vapours of iodine are seen. Following reactions take place :

$MnO_2 + 4HCl\ (conc) \rightarrow MnCl_2 + 2H_2O + Cl_2(g)$

Manganese dioxide   Hydrochloric   Manganese chloride   water

$Cl_2 + 2KBr \rightarrow 2KCl + Br_2$

Chlorine   Potassium bromide   Potassium chloride   (Reddish-brown) Bromine

$2KI + Br_2 \rightarrow 2KBr + I_2$

Potassium iodide   Bromine   Potassium bromide   (Violet) Iodine

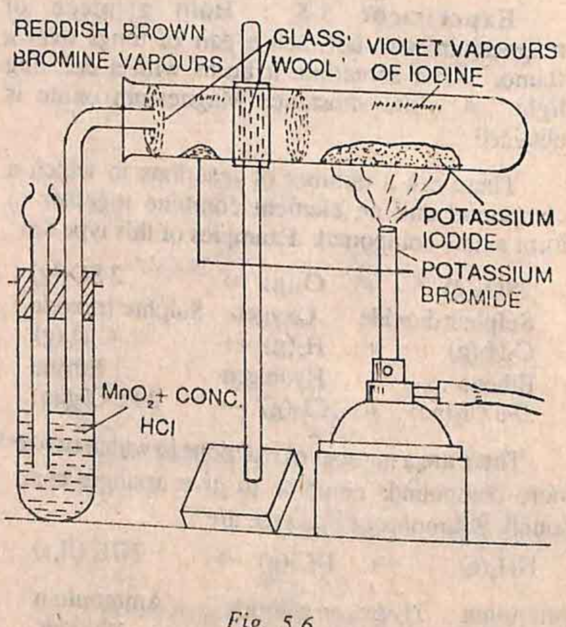


Fig. 5.6

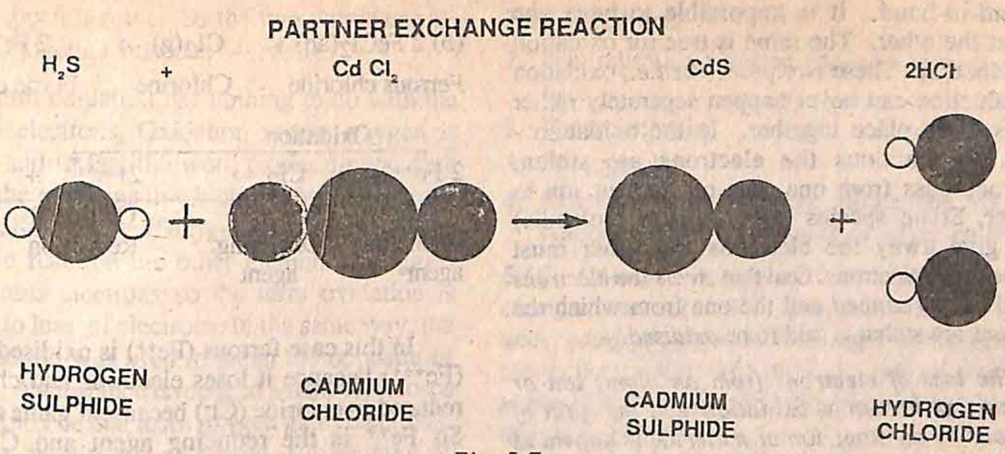


Fig. 5.7

Double decomposition or **Partner exchange reactions :** In these reactions two compounds react to form two new compounds. The cation of one compound displaces the cation of the other compounds. These reactions can be classified as :

Neutralisation Reactions : It is the reaction between acid and base to give salt and water.

$NaOH(aq) + HCl(aq) \rightarrow NaCl + H_2O$

Sodium hydroxide	Hydrochloric acid	Sodium chloride	Water

$Ca(OH)_2(aq) + H_2SO_4(aq) \rightarrow CaSO_4(aq) + 2H_2O$

Calcium hydroxide	Sulphuric acid	Calcium sulphate	Water

$KOH(aq) + HCl(aq) \rightarrow KCl + H_2O$

Potassium hydroxide	Hydrochloric acid	Potassium chloride	Water

Hydrolysis : In these reactions salts like ferric chloride, aluminium chloride solutions etc. when added to boiling water forms insoluble precipitates of hydroxides. As a result of these the medium (acidic or alkaline) of the solution changes.

$FeCl_3(aq) + 3HOH \rightarrow Fe(OH)_3(s) + 3HCl$

Ferric chloride	Water	Ferric hydroxide	Hydrogen-chloride

$AlCl_3(aq) + 3HOH \rightarrow Al(OH)_3(s) + 3HCl$

Aluminium chloride	Water	Aluminium hydroxide	Hydrogen chloride

Precipitation Reactions : In these type of reactions one of the substance formed during the reaction is insoluble in water.

$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$

Barium chloride	Sodium sulphate	Barium sulphate	Sodium chloride

$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

Silver nitrate	Sodium chloride	Silver chloride	Sodium nitrate

Activity : Take silver nitrate solution in a test tube. To it add sodium chloride solution. White precipitates of silver chloride will be formed.

5.4 OXIDATION - REDUCTION REACTIONS :

So far we have studied the reactions that are based on the mode of combination. Now we shall study a set of reactions that are entirely different from the reactions that we have already studied and may include the previous reactions. These reactions are based on the transference of electrons involved in a chemical reactions and are called *oxidation reduction reactions.*

A thief enters into the house of a person to steal some ornaments. In the morning when the owner wakes up, he finds that the ornaments are missing. In this case two processes have happened. The thief has taken away the ornaments and the owner is bewailing their loss. The two processes that are involved in this case are gaining and losing. The thief cannot have the ornaments without their owner losing them and the owner cannot lose them without their falling into the hands of another person. Both gaining and losing go hand-in-hand. It is impossible to have one without the other. The same is true for oxidation and reduction. These two processes *i.e.*, oxidation and reduction can never happen separately rather always takes place together. In the oxidation - reduction reactions the electrons are stolen. Electrons pass from one type of atom or ion to another. Some species (atom, ion or molecule) must give away the electrons and other must receive these electrons. One that *steals* the *electrons* is said to be *reduced* and the one from which the electrons are stolen is said to be *oxidised.*

The loss of electrons from an atom; ion or molecule is known as *oxidation* and *the gain of electrons by an atom, ion or molecule* is known as *Reduction.* The substance that loses electrons is said to be *oxidised* and the one that *gains* electrons is said to be *reduced.* Let us study a few examples :

(*a*) $Zn_{(S)}$+ Cu $SO_4(aq)\rightarrow$ $ZnSO_4(aq)$ + Cu (s)

Zinc Copper Sulphate Zinc Sulphate Copper

Oxidation

$Zn + Cu^{++} \rightarrow Zn^{++} + Cu$

Reduction

In the above case zinc loses electrons and gets oxidised to Zn^{++} and copper ions (Cu^{++}) gain electrons and gets reduced to copper. In this case oxidation of zinc to Zn^{++} is carried by copper ions. So Cu^{++} is the *oxidising agent* and reduction of Cu^{++} to copper is carried by zinc. So zinc is the *reducing agent.*

An oxidising agent is one that brings *about oxidation and reducing agent is one that brings about reduction* —

(*b*) $2\ FeCl_2(aq)$ + $Cl_2(g) \rightarrow$ $2\ FeCl_3(aq)$

Ferrous chloride Chlorine Ferric chloride

Oxidation

$2\ Fe^{++} + Cl_2 \rightarrow 2Fe^{+++} + 2\ Cl^-$

Oxidisied

Reducing agent Oxidising agent Reduction Reduced

In this case ferrous (Fe^{++}) is oxidised to ferric (Fe^{+++}) because it loses electrons and chlorine is reduced to chloride (Cl^-) because it gains electrons. So Fe^{++} is the reducing agent and Cl_2 is the oxidising agent.

(*c*) $Cu(s)+2Ag\ NO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(S)$

Copper Silver nitrate Copper nitrate Silver

Oxidation

$Cu + 2Ag+ \rightarrow Cu^{++} + 2Ag$

Oxidising agent Oxidised Reduced

Reducing agent Reduction

(*c*) $Zn + 2\ H\ Cl(aq) \rightarrow ZnCl_2\ (aq) + H_2$

Zinc Hydrogenchloride Zinc chloride

Oxidation

$Zn + 2H^+ \rightarrow Zn^{++} + H_2$

Oxidised

Reducing agent Oxidising agent Reduction Reduced

Losing and gaining go together

$Fe^{++} + Cl$

$Fe^{++} + Cl$

$\longrightarrow$

Fe^{+++} Cl^-

+

Fe^{+++} Cl^-

Ferrous ions (Reducing agent) Chlorine molecule (Oxidising agent) Ferric ions (Oxidised) Chloride ions (Reduced)

In this reaction chlorine steals away electrons from ferrous ions. So chlorine is the oxidising agent. It is the ferrous ions that has lost an electron and becomes a ferric ion. Chlorine the oxidising agent gets reduced. Ferrous is the reducing agent and gets oxidised to ferric

So from the above reactions we can conclude that oxidation does not occur unless an oxidising agent is present and in the same way reduction does not occur unless a reducing agent is present. In other words oxidation and reduction are complementary processes and one cannot occur unless the other occur simultaneously. One reactant gets oxidised and at the same time the other gets reduced. Because oxidation and reduction always happen together, so the two words can be combined always happen together, so the two words can be combined together into one — *redox*.

The term oxidation has nothing to do with the losing of electrons. Oxidation means oxygen is involved and in fact the word oxidation was first used for the reactions in which a substance gained electrons. So *gaining of oxygen is oxidation*. Since during the reaction the other element that gains oxygen loses electrons so the term oxidation is extended to loss of electrons. In the same way, the term reduction was used to mean the opposite of oxidation. A substance is reduced whenever it loses oxygen. And one that loses oxygen gains electrons. Hence the definition was extended to gaining of electrons.

5.4A. Slow and fast reactions: The rates of chemical reactions vary from the slowest, where the reaction time is of the order of millions of years, to the fastest in which the reaction time is of the order of a few picoseconds. The rate of chemical reactions tells us about the speed of the reaction. Rusting of iron and burning of fuel are both oxidation reactions. Rusting of iron is a typical example of a slow reaction whereas the burning of fuel is an example of a fast reaction. The reaction between dry hydrogen and dry chlorine gases, in the absence of light is an example of an extremely slow reaction but the same reaction becomes violent when the traces of sodium vapours are introduced in the jar of H_2 and Cl_2. For a reaction to occur, molecules must collide with each other. These colliding molecules should possess a minimum energy called threshold energy, so that when they collide, the collision results in a chemical reaction. It has been observed that in the case of slow reactions the value of this threshold energy is very high whereas in the case of fast reaction the value of this threshold energy is very low. So, lower the value of threshold energy the greater will be the reaction rate and depending on this the reactions are classified as slow and fast reactions. A few examples are:

Slow Reactions:

1. Rusting of iron
2. Reaction of copper with air to form basic copper carbonate
3. Formation of fuels
4. Curdling of milk.

Fast Reactions:

1. Neutralisation reaction between an acid and a base
2. Photosynthesis
3. Photography
4. Reaction between silver nitrate and sodium chloride to give precipitate of silver chloride.

5.4B. Catalyst: *A catalyst is a substance which changes the rate of a reaction without itself undergoing any permanent change.* In most of the cases, the catalyst increases the rate of reaction although sometimes a catalyst may even decrease the rate of reaction.

Experimental proof of the effect of a catalyst. Take two dry boiling tubes *A* and *B*. Put about 1g of potassium chlorate in tube *A* and the same amount of potassium chlorate mixed with a little manganese dioxide in tube *B*. Now assemble the two sets of apparatus as shown in Fig. 5.7A and heat the two tubes.

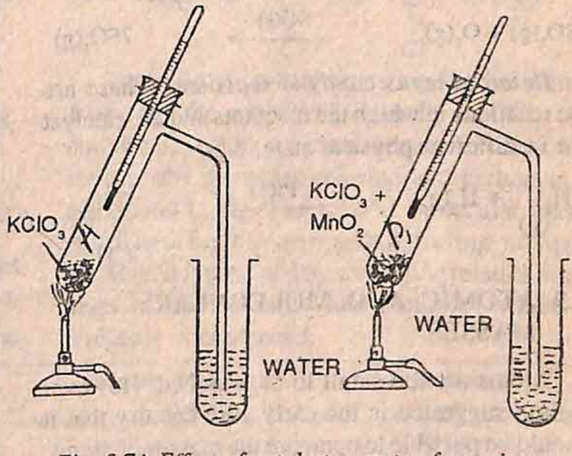


Fig. 5.7A Effect of catalyst on rate of reaction

It is observed that there is hardly any liberation of oxygen in tube *A*, which contains only potassium chlorate, even on heating to a high temperature. On the other hand, in tube *B*, which contains manganese dioxide mixed with potassium chlorate, oxygen is liberated rapidly even at a lower temperature.

$$2KClO_3(s) \xrightarrow{\text{Heat}} 2KCl(s) + 3O_2(g)$$

(No appreciable change)

$$2KClO_3(s) \xrightarrow{MnO_2;\ \text{heat}} 2KCl(s) + 3O_2(g)$$

After the reaction is over, manganese dioxide is found to be left unchanged.

Sometimes a catalyst may be added to decrease the rate of an undesirable reaction. For example, addition of some alcohol decreases the rate of decomposition of hydrogen peroxide. Such a catalyst is generally called *negative catalyst*.

Catalysts are used in the manufacture of many important chemicals. For example, vanadium pentoxide is used as a catalyst in the manufacture of sulphuric acid by contact process, iron oxide is used as a catalyst in the Haber's process for the manufacture of ammonia and finely divided nickel is used as a catalyst in the hydrogenation of oils to form *vegetable ghee*.

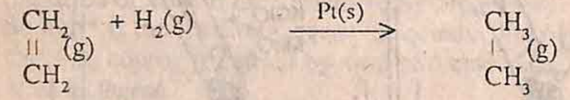
All the catalytic reactions can be classified under the following two heads:

Homogeneous catalytic reactions: These are the reactions in which the catalyst and the reactants are in the same physical state, e.g., hydrolysis of ester in the presence of dil. hydrochloric acid.

$$CH_3COOC_2H_5(ag) + H_2O(P) \xrightarrow{HCl(ag)} CH_3COOH(ag) + C_2H_5OH(ag)$$

$$2SO_2(g) + O_2(g) \xrightarrow{NO(g)} 2SO_3(g)$$

Heterogeneous catalytic reactions: These are the reactions in which the reactants and the catalyst are in different physical state, e.g.,



$$\begin{array}{l}CH_2\\ \| \ (g)\\ CH_2\end{array} + H_2(g) \xrightarrow{Pt(s)} \begin{array}{l}CH_3\\ | \ (g)\\ CH_3\end{array}$$

5.5 ATOMIC AND MOLECULAR MASSES

Atoms are too small to be weighed. However Dalton suggested in the early 19th century that it should be possible to compare the masses of atoms, although the atoms themselves could not be weighed. Chemists use *a relative atomic mass scale* to compare the masses of different atoms. In this scheme — at first the element hydrogen was chosen as the standard against which the masses of other atoms were compared. Hydrogen was assigned the arbitrary value of 1. Hydrogen was chosen initially because it had the smallest atoms which could be easily given a relative atomic mass of one. Later on when scientists realised that an element contains atoms of different masses (*i.e.*, isotopes) it became necessary to choose a single isotope as the reference for relative atomic mass.

In 1961 the isotope of carbon having six protons and six neutrons was chosen as the new standard for determining the relative atomic mass.

Why an isotope of carbon (carbon $^{-12}$ has been chosen as a standard for determining the relative atomic masses of the elements rather than an isotope of hydrogen or oxygen ? This is because

(a) carbon is a very common element.

(b) It is solid at room temperature and pressure, so easy to store and transport than gas.

(c) It forms stable components with a large number of elements.

On the carbon^{-12} scale, the atoms of the isotope of carbon (^{12}C) is assigned a relative atomic mass of 12 and the relative atomic masses of all other atoms are obtained by comparing with the mass of ^{12}C atom. *One atomic mass unit* (a.m.u) is equal to one twelfth of the mass of one carbon^{-12} atom. Hence the *atomic mass of* an element is the relative mass of one atom of the element as compared with an atom of carbon^{-12} as the standard. One atom of hydrogen has one-twelfth the mass of the standard carbon^{-12}. So atomic mass of hydrogen is 1 a.m.u. The relative atomic mass of magnesium on the C^{-12} scale is 24.312. It means the average mass of Mg atom and the mass of carbon atom are in the ratio 24.312 : 12.000 or, an atom of magnesium is nearly twice as heavy as an atom of carbon^{-12}. A few relative atomic masses are listed in the table below:

Formula weight and Molecular weight or Formula mass and Molecular mass: The molecular mass is the relative mass of a molecule of a substance as compared to the atomic mass of C^{-12} taken as 12 a.m.u., i.e., how many times a molecule of a substance is heavier than one-twelfth the mass of an atom of carbon taken as 12 a.m.u.

For example : Molecular mass of oxygen is 32 a.m.u. it means a molecule of oxygen is 32 times heavier than one-twelfth the mass of an atom of carbon^{-12}. The sum of the atomic masses of the atoms in the formula of a molecule is its *molecular mass*. **For example :**

Formula of sulphuric acid is H_2SO_4

So its molecular mass will be :

=2 x Atomic mass of Hydrogen + Atomic mass of Sulphur + 4 x atomic mass of oxygen.

$=2 \times 1 + 32 + 4 \times 16 = 2 + 32 + 64 = 98$ a.m.u.

Hence one molecule of sulphuric acid is 98/12 times as heavy as carbon^{-12}

When the true formula of a molecule of a substance is known, the molecular mass can be calculated by adding the atomic masses of the atoms given in the formula of the substance. It is correct to speak of the molecular mass of substances that consist of discrete molecules such as

Carbon dioxide (CO_2) Ethane (C_2H_6)
Water (H_2O)
Ammonia (NH_3) Sulphur dioxide (SO_2), etc.

But ionic compounds such as potassium chloride (KCl) sodium chloride (NaCl), sodium hydroxide (NaOH) do not consist of physically distinct and electrically neutral molecules either in the solid state or in the gaseous state or in solution. So the term *formula mass* is used to represent the composition of such substances. **For example :**

Formula mass of sodium chloride is :

= Atomic mass of sodium + Atomic mass of chlorine
= 23 + 35.5
= 58.5 a.m.u.

Formula mass of potassium nitrate (KNO_3)

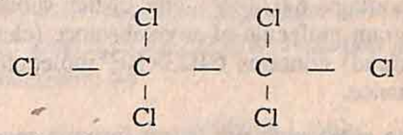
= Atomic mass of potassium + Atomic mass of nitrogen+3xAtomic mass of oxygen.
= 39 + 14 + 3 x 16
= 39 + 14 + 48
= 101 a.m.u.

For substances existing as discrete molecules the molecular mass and formula mass are the same.

Empirical and molecular formulas : The chemical formulae K + Cl and CH_4 represent two different types of structures; but both show the ratio of the particles of the combining elements in their structures. A formula that shows the ratio of the number of the respective atoms of ions in a compound is known as empirical formula. The formulae for all electrovalent compounds (ionic compounds) are empirical formulae because they show the simplest ratios of the ions in the crystal lattice.

Methane is composed of separate molecules and the formula CH_4 is known as molecular formula. It shows the actual number and types of atoms in a molecule of that compound. In this case both empirical formula and molecular formula are the same.

The molecular formula of hexachloro-ethane is C_2Cl_6. In each molecule of C_2Cl_6 the two carbon atoms are covalently bonded to each other and also to six chlorine atoms.



The ratio of carbon atoms to chlorine atoms is 1 : 3. So the empirical formula is CCl_3. This empirical formula indicates the simplest combining ratio of the elements chlorine and carbon in the compound C_2Cl_6. The relative molecular mass is calculated for the compounds having molecular formula. It is obtained by adding the relative atomic mass of all the individual atoms of which the molecule is composed.

Substance	Molecular Formula	Relative Molecular Mass	Empirical Formula
Potassium chloride			K^+Cl^-
Sodium nitrate			Na^+NO_3
Magnesium carbonate			$Mg^{2+}CO_3^{2}$
Nitric acid			H^+NO_3
Benzene	C_6H_6	6 x 12 + 6 x 1 = 78	CH
Methane	CH_4	12 + 4 x 1 = 16	CH_4
Acetylene	C_2H_2	12 x 2 + 2 x 1 = 26	CH
Glucose	$C_6H_{12}O_6$	12 x 6 + 1 x 12 + 16 x 6 = 180	CH_2O

5.6 GRAM ATOMIC MASS AND GRAM MOLECULAR MASS

We have studied the atomic masses of the elements relative to the mass of carbon^{-12} isotope taken as 12 a.m.u. Relative atomic mass is the mass of an atom of the element compared with one-twelfth the mass of one atom of C^{-12}. In case we use gram as the unit to express these relative atomic masses we have *gram atom.* So one gram atom of any element is the relative atomic mass of that element expressed in grams. **For example :** the relative atomic mass of copper is 63.5 a.m.u.

gram atomic mass of sulphur will be 32·1 gram.

Gram molecular mass is the molecular mass of a substance expressed in grams. Gram molecular mass of water is 18 gram.

Mole : Mole is a chemists unit that represents a set of $6{\cdot}023 \times 10^{23}$ particles. These can be atom, ion or molecule or any other species. For **example** : 1 mole sodium atom means $6{\cdot}023\times10^{23}$ atoms of sodium. 1 mole oxygen molecule means $6{\cdot}023\times10^{23}$ molecules of oxygen. 1 mole water means $6{\cdot}023\times10^{23}$ molecules of water. 1 mole sodium ions means $6{\cdot}023\times10^{23}$ ions of soduim. 1 mole sodium chloride contains $6{\cdot}023\times10^{23}$ sodium ions and $6{\cdot}023\times10^{23}$ chloride ions.

This number $6{\cdot}023\times10^{23}$ is known as *Avogadro's Number* or *Avogadro's constant* in the honour of Amaedo Avogadro.

It has been found that one gram atom of any element contains $6{\cdot}023\times^{23}$ atoms of that substance and one gram molecule of any substance (element of compound) contains $6{\cdot}023\times10^{23}$ molecules of that substance.

In the same way one gram formula mass of any substance will contain $6{\cdot}023\times10^{23}$ ions of each type.

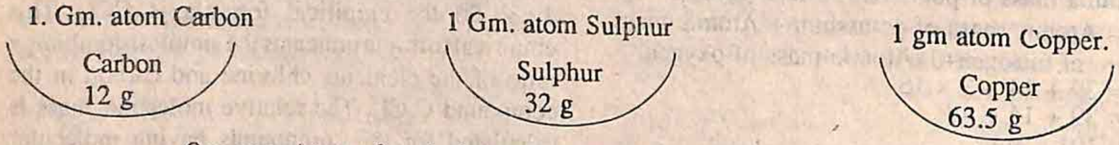
So we can conclude that *moles give us a relationship between number and amount.* Hence one mole of any element will contain the same number of atoms.

And one mole of any compound will contain the same number of molecules.

Hence 1 mole of Carbon dioxide means
$= 6{\cdot}023 \times 10^{23}$ molecules of CO_2
$= 12+16\times2 = 44$ gms of CO_2
= one gram molecular mass of CO_2

1 mole of Sodium atom
= 1 gm atom of sodium
$= 6{\cdot}023 \times 10^{23}$ atoms of sodium
= 23 gram of sodium

1 mole of Sodium chloride
= 1 gram of formula mass of sodium chloride
$= 23 + 35{\cdot}5 = 58{\cdot}5$ gram of soduim chloride
$= 6{\cdot}023 \times 10^{23}$ sodium chloride species
$= 6{\cdot}023 \times 10^{23}$ sodium ions (Na^+) and
6.023×10^{23} chloride ions (Cl^-)



One gram atom of each element contains the same number of atoms.
Mole is abbreviated as mol.

So Number of moles of a substance

$$= \frac{\text{Mass of a substance in grams}}{\text{Mass of 1 mole of a substance}}$$

No.of mole atoms

$$= \frac{\text{Weight in grams}}{\text{gram atomic mass}}$$

No. of mole molecules

$$= \frac{\text{Weight in grams}}{\text{gram molecular mass}}$$

No. of moles

$$= \frac{\text{Number of Particles}}{6{\cdot}023\times10^{23}}$$

No. of mole atoms

$$= \frac{\text{No. of atoms}}{6{\cdot}023\times10^{23}}$$

No. of mole molecules

$$= \frac{\text{No. of molecules}}{6{\cdot}023\times10^{23}}$$

Example : Calculate the number of atoms in $1{\cdot}792 \times 10^{-12}$ gram of silver. (Atomic mass = 107·87 a.m.u.)

Solution :

Since atomic mass of silver = 107·87 a.m.u.
So gram atomic mass of silver = 107·87 gram
1 mole silver atom weigh 107·87 gram
$6{\cdot}023\times10^{23}$ atoms of silver weigh = 107·87 grams

So 107.87 grams of silver contain 6.023×10^{23} atoms of Ag

1 gram of silver contains $\frac{6{\cdot}023\times10^{23}}{107.87}$ atoms of Ag

$1{\cdot}792\times10^{-12}$ gram of silver contains $= \dfrac{6{\cdot}023\times10^{23}}{107.87}$

$$\times\, 1{\cdot}792\times10^{-12}$$

$$= \frac{6{\cdot}023\times10^{23}}{107870} \times 1{\cdot}792\times10^{-12}$$

$$= 10^{10} \text{ atoms of silver.}$$

Example 5.2 : Calculate the number of carbon atom in 6 grams of carbon (Atomic mass = 12 a.m.u).

Solution : Atomic mass of carbon 12 a.m.u.

So 1 mole of carbon weighs 12 grams

Hence 12 grams of carbon contains $6{\cdot}023\times10^{23}$ carbon atoms

1 gram of carbon contains

$= \dfrac{6{\cdot}023\times10^{23}}{12}$ carbon atoms.

6 grams of carbon contains

$= \dfrac{6{\cdot}023\times10^{23}}{12} \times 6$ carbon atoms

$= 3{\cdot}0115\times10^{23}$ carbon atoms. **Ans.**

Example 5.3 : How many moles of NH_3 are there in 3.4 grams of ammonia ? (Atomic mass N=14 a.m.u)
H=1 a.m.u)

Solution : Molecular mass of NH_3 = Atomic mass of Nitrogen

$+3 \times$ Atomic mass of H atom

$= 14+3\times 1$

$= 14+3 = 17$ a.m.u.

So 1 mole of Ammonia weighs = 17 grams
Hence 17 grams of Ammonia contains

$= 6{\cdot}023\times10^{23}\ NH_3$

1 gram of Ammonia contains

$= \dfrac{6{\cdot}023\times10^{23}}{17}\ NH_3$

3·4 grams of Ammonia contains

$= \dfrac{6{\cdot}023\times10^{23}}{17} \times \dfrac{3{\cdot}4}{10}$

$= 12{\cdot}046 \times 10^{22}\ NH_3$ **Ans.**

Example 5.4 : How many atoms are there in 18 gm of water ? (Atomic mass of H = 1 a.m.u ; O =16 a.m.u)

Solution : Molecular formula of water = H_2O

Molecular mass of water = $1 \times 2 + 16$

= 18 a.m.u

So 1 mole of water weighs = 18 grams.

Hence 18 grams of water contains 1 mole water molecules.

1 mole water contains

= 2 mole Hydrogen atoms and 1 mole Oxygen atom

2 mole Hydrogen atoms

$= 2 \times 6{\cdot}023 \times 10^{23}$ atoms of Hydrogen **Ans.**

1 mole Oxygen atom

$= 6{\cdot}023 \times 10^{23}$ atoms of Oxygen **Ans.**

If the substance is a gas then one mole of any gas at S.T.P occupies 22·4 litre volume. By S.T.P. we mean standard temperature and pressure. Standard temperature is 273 K and standard pressure is 1 atmosphere. So

1 mole of any gas at S.T.P

contains 1 mole molecules *i.e.*, 6.023×10^{23} molecules

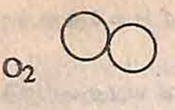
weigh equal to gram molecular mass

occupy 22·4 lit volume at 273 K and 1 atmospheric pressure

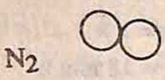
Derivation of Formula : Following steps are kept in mind while deriving the formula of a new compound:

1. Determine the percentage composition of the compound :

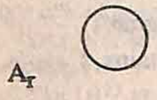
The percentage composition can be determined experimentally by analysis or synthesis. A weighed sample of the compound is separated into its constituent elements and the weight of each element in the sample is determined. **For example** : 1·0000 gram sample of a certain compound is found to have 0·2729 gram of carbon and 0·7271 gram of oxygen. Then :



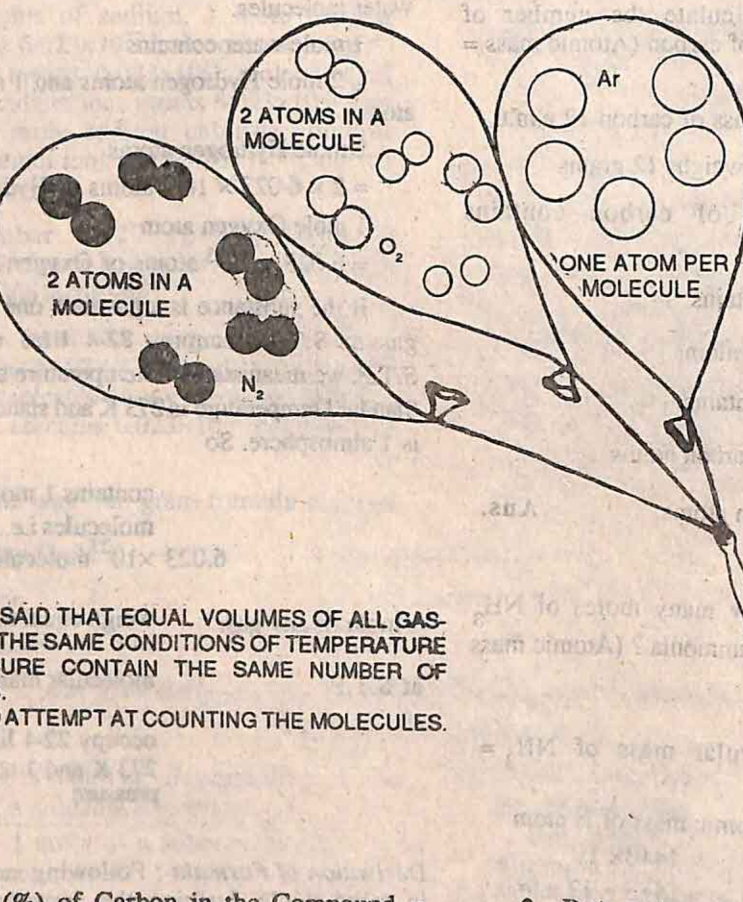
THE MOLECULAR WEIGHT OF OXYGEN IS 32. A GRAM MOLECULAR WEIGHT OF OXYGEN IS 32 g. IN 32 g. OF OXYGEN THERE WILL BE 6·023 × 10^{23} MOLECULES.



THE MOLECULAR WEIGHT OF NITROGEN IS 28. A GRAM MOLECULAR WEIGHT OF NITROGEN IS 28 g. IN 28 g. OF NITROGEN THERE WILL BE 6·023 × 10^{23} MOLECULES.



THE MOLECULAR WEIGHT OF ARGON IS 40. A GRAM MOLECULAR WEIGHT OF ARGON IS 40 g. IN 40 g. OF ARGON THERE WILL BE 6·023 × 10^{23} MOLECULES.



AVOGADRO SAID THAT EQUAL VOLUMES OF ALL GASES UNDER THE SAME CONDITIONS OF TEMPERATURE AND PRESSURE CONTAIN THE SAME NUMBER OF MOLECULES.
HE MADE NO ATTEMPT AT COUNTING THE MOLECULES.

Percentage (%) of Carbon in the Compound = $\frac{0{\cdot}2729 \text{ gm}}{1{\cdot}0000 \text{ gm}} \times 100 = 27{\cdot}29\%$

% of Oxygen in the Compound

$$= \frac{0{\cdot}7271 \text{ gm}}{1{\cdot}0000 \text{ gm}} \times 100$$

$$= 72{\cdot}71\ \%$$

2. Determine the relative number of atoms of the elements in the compound by dividing the percentage of elements calculated above by the relative atomic masses of the respective atoms.

Element	*%*	*Atomic mass*	*Relative No. of atoms*	*Divided by the common divider*	*Smallest No. of atom*
C	27·29	12·00	$\frac{27{\cdot}29}{12} = 2{\cdot}27$	$\frac{2{\cdot}27}{2{\cdot}27} = 1$	1
O	72·71	16·00	$\frac{72{\cdot}71}{16} = 4{\cdot}54$	$\frac{4{\cdot}54}{2{\cdot}27} = 2$	2

3. Divide the relative number of atoms by the common divider : Since only whole number of atoms can enter into chemical combination so the relative number of atoms are reduced to whole numbers of atoms by dividing with the smallest number. So in the above example there are twice as many oxygen atoms as carbon atoms in the compound.

So the simplest formula for the compound must be CO_2.

The simplest formula that indicates the ratio of the numbers of different kinds of atoms present in the compound is known as *empirical formula* and the formula that indicates the actual number of each kind of atoms present per molecule is known as *Molecular formula* .

In order to determine the molecular formula of the compound the molecular mass of the compound is divided by the empirical formula weight. In case the number comes out to be one it means empirical formula is the same as the molecular formula. If the value comes out to be two then the number of each kind of atoms in the molecular formula are twice the number of each kind of atoms in the empirical formula and so on.

Example 5.5 : An oxide of sulphur contains 50% sulphur and 50% oxygen. If the relative molecular mass of the compound is 64 and relative atomic masses of sulphur and oxygen are 32 and 16 respectively. Calculate the empirical formula and molecular formula of the oxide.

Solution :

Element	*%*	*Atomic mass*	*Relative No. of Atoms*	*Divided by common divider*	*Smallest No. of atoms*
Sulphur	50	32	$\frac{50}{32}$ = 1·56	$\frac{1.56}{1.56}$ =1	1
Oxygen	50	16	$\frac{50}{16}$ = 3·12	$\frac{3.12}{1.56}$ =2	2

So one atom of Sulphur combines with two atoms of Oxygen.

Hence empirical formula is SO_2.

Empirical formula mass = Atomic mass of Sulphur + 2 × At mass of Oxygen

$= 32 + 16 \times 2$

$= 32 + 32 = 64$ a.m.u.

Relative molecular mass = 64.

Since the empirical formula mass is equal to the relative molecular mass so the molecular formula is also SO_2.

Example 5.6 : 40 gram of a compound contains 12·8 gram of carbon ; 1·6 gram of hydrogen and 25·6 gram of oxygen. If the relative molecular mass of the compound is 150 and the relative atomic mass of carbon, hydrogen and oxygen are 12, 1 and 6 respectively, then calculate the empirical and molecular formula of the compound.

Solution : Weight of the Compound = 40 grams

Weight of Carbon = 12·8 grams
Weight of Hydrogen = 1·6 grams
Weight of Oxygen = 25·6 grams

$$\% \text{ of Carbon} = \frac{\text{weight of Carbon}}{\text{weight of the compound}} \times 100$$

$$= \frac{12.8}{40} \times 100 = 32\%$$

$$\% \text{of Hydrogen} = \frac{\text{weight of Hydrogen}}{\text{weight of the compound}} \times 100$$

$$= \frac{1.6}{40} \times 100 = 4\%$$

$$\% \text{ of Oxygen} = \frac{\text{weight of Oxygen}}{\text{weight of the compound}} \times 100$$

$$= \frac{25.6}{40} \times 100 = 64\%$$

Element	*%*	*Atomic mass*	*Relative No. of Atoms*	*Divided by common divider*	*Simple whole Number*
C	32	12	32/12 = 2·67	$\frac{2.67}{2.67}$ = 1	2
H	4	1	4/1 = 4·00	$\frac{4.0}{2.67}$ = 1·5	3
O	64	16	64/16 = 4·00	$\frac{4.0}{2.67}$ = 1·5	3

The above calculations indicate that the ratio of Carbon to Hydrogen to Oxygen atoms in the compound is :

1 Carbon : 1·5 Hydrogen : 1·5 Oxygen

1 C : 1·5 H : 1·5 Oxygen

But it is not possible to have a fraction of an atom. So the empirical formula must be a whole number ratio.

This simplest whole number ratio is = 2C : 3H : 3O

Hence empirical formula is $C_2H_3O_3$

Empirical formula mass of the compound is

= 2 × Atomic mass of C + 3 × Atomic mass of H + 3 × Atomic mass of O

$= 2 \times 12 + 3 \times 1 + 3 \times 16$

$= 24 + 3 + 48$

$= 75$

Relative molecular mass = 150

$$n = \frac{\text{Relative molecular mass}}{\text{Empirical formula mass}} = \frac{150}{75} = 2$$

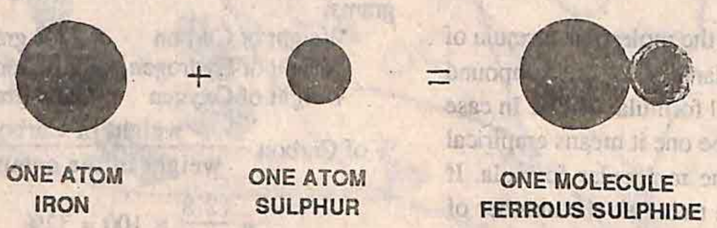
So molecular formula = Empirical formula × n

$= C_2H_3O_3 \times 2$

$= C_4H_6O_6$

BALANCING EQUATIONS

1.



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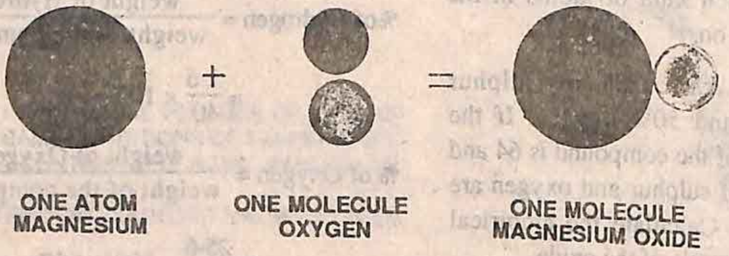

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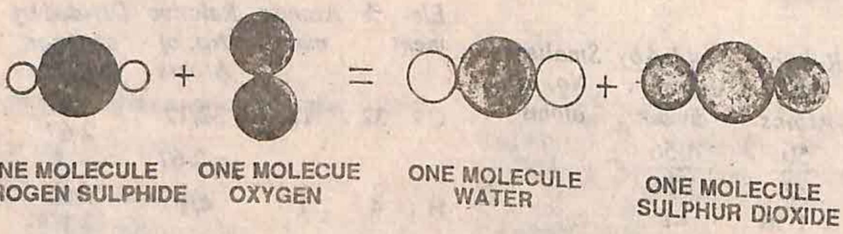

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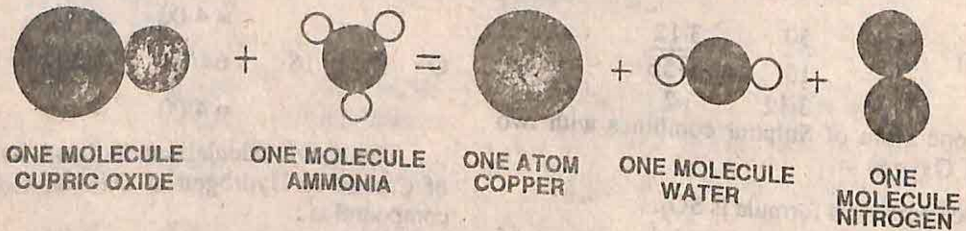


Chemical Equations : A chemical equation summarises the quantitative investigations into a chemical reaction. It also indicates the member of moles of reactants used in the reaction and the number of moles of products formed. A chemical equation represents a chemical change by means of symbols and formulas. The following procedure is used for writing the chemical equation.

1. First of all write down the formulas of all the reactants and the products in the equation. The formula of the reactants is written on the left-hand side and the formula of the products is written on the right-hand side of the arrow. These formulas are constant and cannot be changed. **For example :** The formula of methane is CH_4 so in the chemical equation the formula of methane must always be written as CH_4 or multiples of CH_4 such as 2 CH_4 or 3 CH_4 or 4 CH_4 and so on.

Zn	+	2 HCl ⟶	Zn Cl_2 +	H_2
Zinc	+	Hydrochloric Acid	Zinc Chloride	Hydrogen
Reactants			Products	

2. After this subscript is written within brackets before each reactants and the products to indicate whether the substance is a solid, gas or liquid or dissolved in water.

   For solid (s) is written,
   liquid (l) is written,
   gas (g) is written,
   dissolved in water (aq) is written.

3. Now count the number of particles (atoms or ions) of each element on either side of the equation and make the number and kind of particles equal on both sides of the equation without altering the formula for any substance. The smallest number for each substance is used to balance the equation.

For example : Let us balance the equation for burning of Hydrogen in Oxygen.

(*a*) Hydrogen burns in Oxygen to give water

Step (*i*) $H_2 + O_2 \longrightarrow H_2O$

Step (*ii*) $H_2\,(g) + O_2\,(g) \longrightarrow H_2O\,(g)$

Step (*iii*) No. of atoms on the left side = Two Hydrogen atoms and two Oxygen atoms

No. of atoms on the right side

= Two Hydrogen atoms and one Oxygen atom

Step (*iv*) The equation is not balanced, so in order to balance the Oxygen the formula for water (H_2O) is multiplied by two

$$H_2\,(g) + O_2\,(g) \longrightarrow 2\,H_2O\,(g)$$

Now the reactant H_2 must be multiplied 2 in order to balance this.

So the balanced equation for this reaction is $2\,H_2\,(g) + O_2\,(g) \longrightarrow 2\,H_2O\,(g)$.

Example : Write and balance the equation for the burning of ethane in air.

Solution : When ethane gas burns in air it reacts with oxygen of the air giving heat, carbondioxide and water vapour.

Step (*i*) $C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$

Step (*ii*) Make it more informative.

$$C_2H_6\,(g) + O_2\,(g) \longrightarrow CO_2\,(g) + H_2O\,(g)$$

Step (*iii*) No. of particles (atoms) on the left side = 2 carbon ; 6 hydrogen and 2 oxygen.

No. of particles (atoms) on the right side = 1 carbon; 3 oxygen and 2 hydrogen.

Step (*iv*) The equation is not balanced. (Unbalanced equation is known as Skeleton equation).

Two carbon and six hydrogen atoms would produce two carbon dioxide and three water molecules.

$$C_2H_6\,(g) + O_2\,(g) \longrightarrow 2\,CO_2\,(g) + 3\,H_2O\,(g)$$

In order to have whole number of oxygen molecules water is multiplied by six instead of three.

$$C_2H_6\,(g) + O_2\,(g) \longrightarrow 2CO_2\,(g) + 6\,H_2O\,(g)$$

Now to balance hydrogen multiply C_2H_6 by two and to balance carbon multiply CO_2 by four $2\,C_2H_6\,(g) + O_2\,(g) \longrightarrow 4\,CO_2\,(g) + H_2O\,(g)$.

Now the equation is not balance w.r.t. oxygen. So multiply oxygen by 7

$$2\,C_2H_6\,(g) + 7\,O_2\,(g) \rightarrow 4\,CO_2\,(g) + 6\,H_2O\,(g).$$

The above method for the balancing of the chemical equations is known as Hit & Trial Method. There is another method for balancing the chemical equations. It is known as *Partial Equation Method*. This method for balancing the equations is used for the reactions that take place in a number of steps. **For example :** Consider the reaction between copper and conc. nitric acid.

$$\underset{\text{(Conc.)}}{Cu + HNO_3} \longrightarrow Cu\,(NO_3)_2 + H_2O + NO_2$$

(*a*) Write partial equations for the various steps

$$HNO_3 \longrightarrow H_2O + NO_2 + O$$
$$Cu + O \longrightarrow CuO$$

$$CuO + HNO_3 \longrightarrow Cu(NO_3)_2 + H_2O$$

(*b*) Balance each slip by hit and trial method

$$\begin{aligned} 2\,HNO_3 &\longrightarrow H_2O + 2\,NO_2 + O \\ Cu + O &\longrightarrow Cu\,O \\ CuO + 2\,HNO_3 &\longrightarrow Cu(NO_3)\,2 + H_2O \end{aligned}$$

(*c*) The intermediate products that do not appear in the final equation are cancelled out. And the partial equations are then added to get the final balanced equation.

So the above equations can be rewritten as :

$$\begin{array}{rcl} 2\,HNO_3 & \longrightarrow & H_2O + 2\,NO_2 + O \\ Cu + O & \longrightarrow & CuO \\ CuO + 2\,HNO_3 & \longrightarrow & Cu(NO_3)_2 + H_2O \\ \hline Cu + 4HNO_3 & \longrightarrow & Cu(NO_3)_2 + 2H_2O + 2NO_2 \end{array}$$

Reaction between acidified Potassium permanganate and ferrous sulphate

$$KMnO_4 + H_2SO_4 + FeSO_4 \longrightarrow K_2SO_4 + MnSO_4 + H_2O + Fe_2(SO_4)_3.$$

$$\begin{array}{lcl} 2KMnO_4 + 3H_2SO_4 & \longrightarrow & K_2SO_4 + 2\,MnSO_4 + 3H_2O + 5\,(O) \\ [2FeSO_4 + H_2SO_4 + O & \longrightarrow & Fe_2(SO_4)_3 + H_2O] \times 5 \\ \hline 2KMnO_4 + 3H_2SO_4 + 10FeSO_4 + 5H_2SO_4 + 5\,(O) & \longrightarrow & K_2SO_4 + 2MnSO_4 + 3\,H_2O + 5\,(O) + 5Fe_2(SO_4)_3 + 5H_2O \\ \hline 2KMnO_4 + 8H_2SO_4 + 10FeSO_4 & \longrightarrow & K_2SO_4 + 2\,MnSO_4 + 8H_2O + 5Fe_2(SO_4)_3 \\ \hline \end{array}$$

Ionic Equations : There are a number of reactions that take place between ions and not the molecules. Consider the reaction between Zinc and Hydrochloric acid

$$Zn\,(s) + 2\,HCl\,(aq) \longrightarrow ZnCl_2\,(aq) + H_2\,(g)$$

Let us write it in the form of ions that are present in the solution

$$Zn\,(s) + 2\,H^+\,(aq) + 2\,Cl^-\,(aq) \longrightarrow Zn^{2+}\,(aq) + 2\,Cl^-\,(aq) + H_2\,(g)$$

Cancelling the common ions on both the sides we have

$$Zn\,(s) + 2H^+\,(aq) \longrightarrow Zn^{2+}\,(aq) + H_2\,(g)$$

This equation is known as *Ionic equation*. The charges in ionic equations on both the sides must be equal. This is known as balancing of ionic equation. A few more examples are :

(*a*) Reaction between HCl and NaOH

$$H^+\,(aq) + Cl^-\,(aq) + Na^+\,(aq) + OH^-\,(aq) \longrightarrow Na^+(aq) + Cl^-\,(aq) + H^2O\,(l)$$

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l) \textit{ Ionic equation.}$$

(*b*) Reaction between acids and carbonates

$$2\,HCl_{(aq)} + Na_2CO_3\,(aq) \longrightarrow 2NaCl_{(aq)} + H_2O\,(l) + CO_2\,(g)$$

$$2H^+\,(aq) + CO_3^{2-}\,(aq) \longrightarrow CO_2\,(g) + H_2O\,(l)$$

Ionic Equation

(*c*) Double decomposition reactions

$$CuSO_4\,(aq) + 2\,NaOH\,(aq) \longrightarrow Cu\,(OH)_2\,(S) + Na_2SO_4\,(aq)$$

$$Cu^{++}(aq) + 2OH^-(aq) \longrightarrow Cu(OH)_2(S) \text{ Ionic equation}$$

(*d*) Single displacement reactions

(*i*) $Zn\,(s) + CuSO_4\,(aq) \longrightarrow ZnSO_4\,(aq) + Cu\,(s)$

$Zn\,(s) + Cu^{++}\,(aq) \longrightarrow Zn^{++}\,(aq) + Cu\,(s)$

(*ii*) $2\,Al + 3\,H_2SO_4 \longrightarrow Al_2(SO_4)_3 + 3\,H_2$

$Al + 2\,H^+ \longrightarrow Al^{3+} + H_2$

Since the charges on both sides must be equal so

$$2\,Al + 6H^+ \longrightarrow 2\,Al^{3+} + 3\,H_2$$

INFORMATION GIVEN BY CHEMICAL EQUATIONS :

Consider the balanced chemical equation

$$2\,H_2\,(g) + O_2\,(g) \longrightarrow 2\,H_2O\,(g)$$

This equation tells us that two moles of gaseous H_2 reacts with one mole of gaseous oxygen to give two moles of gaseous water.

Any balanced chemical equation gives us the following informations :

1. The name of the reactants and the products.
2. The relative number of moles of each of the reactants and the products
3. The physical state of the reactants and products.

5.7 LIMITATIONS OF BALANCED CHEMICAL EQUATIONS

A balanced chemical equation does not tell us about

1. the conditions necessary for the reaction to occur.
2. concentrations of the reactants.
3. whether or not all the reactants will change into products.
4. mechanism of the reaction.
5. energy changes taking place during the reaction.
6. the reaction rate.

NUMERICALS :

Example 5.7 Calculate the percentage of different elements in Ammonium dichromate

[Atomic mass of N=14 ; Cr= 52 ; O = 16 ; H=1]

Solution : Molecular formula of Ammonium dichromate = $(NH_4)_2 Cr_2 O_7$.

$$\text{Molecular mass of Ammonium dichromate} = \left\{ \begin{array}{l} 2 \times [\text{At.mass of N}+4 \\ \times \text{At.mass of H}] \end{array} \right. + 2 \times \text{At. mass of Cr} + 7 \times \text{At. mass of O}] \}$$

$= [2 \times (14+4) + 2 \times 52 + 7 \times 16)]$

$= [2 \times (14+4) + 104 + 112)]$

$= 36 + 104 + 112 = 252$

% of

$$N = \frac{\text{gm. at. mass of Nitrogen}}{\text{gm. mol. mass of } (NH_4)_2 Cr_2 O_7} \times 100$$

$= \frac{28}{252} \times 100 = 11{\cdot}2$ % **Ans.**

% of Hydrogen

$$= \frac{\text{gm. at. mass of Hydrogen}}{\text{gm mol mass of } (NH_4)_2 Cr_2 O_7} \times 100$$

$= \frac{8}{252} \times 100 = 3{\cdot}2$ % **Ans.**

% of Chromium

$$= \frac{\text{gm. at. mass of chromium}}{\text{gm mol. mass of } (NH_4)_2 Cr_2 O_7} \times 100$$

$= \frac{104}{252} \times 100 = 41{\cdot}2$% **Ans.**

% of Oxygen

$$= \frac{\text{gm. at. mass of Oxygen}}{\text{gm mol. mass of } (NH_4)_2 Cr_2 O_7} \times 100$$

$= \frac{112}{252} \times 100 = 4\ {\cdot}2$ % **Ans.**

Example 5.8 Calculate the percentage of water in $CuSO_4 . 5H_2O$

Solution : Formula weight of $CuSO_4 . 5H_2O$ = At. mass of Cu + At. mass of S + 4 × At. mass of Oxygen

+ 5 [2 × At. mass of Hydrogen + At. mass of Oxygen)

$= 63{\cdot}5 + 32 + 4 \times 16 + 5\,(2 \times 1 + 16)$

$= 63{\cdot}5 + 32 + 64 + 90 = 249{\cdot}5$

So % of water

$$= \frac{\text{Total weight of water in gms}}{\text{gm. formula weight}} \times 100$$

$= \frac{90}{249{\cdot}5} \times 100 = 36{\cdot}1$ %

Example : Calculate the weight of Oxygen in moles and in grams that is produced during the thermal decomposition of 100 gms of potassium chlorate.

Solution : *Chemical Equation*

$$2\,KClO_3\,(s) \xrightarrow{\text{Heat}} 2\,KCl\,(s) + 3\,O_2\,(g)$$

Potassium chloride　　　Potassium chloride　Oxygen

Step 1

First of all calculate the number of moles of $KClO_3$ in 100 gm of $KClO_3$

$$= \frac{100 \text{ gm}}{122{\cdot}6 \text{ gm/mole}}$$

$= 0{\cdot}816$ mole of $KClO_3$

According to the equation

2 moles of $KClO_3$ (s) will give = 3 moles of O_2 (g)

1 mole of $KClO_3$ (s) will give $= \frac{3}{2}$ moles of O_2 (g)

0·816 mole of $KClO_3$ (s) will give

$= \frac{3}{2} \times 0{\cdot}816$ mole of O_2 (g)

$= 1{\cdot}224$ moles of O_2 (g) Ans.

No. of grams of Oxygen produced

= No. of moles × gram molecular of oxygen

= 1·224 moles × 32·0 gm/mole

= 39·2 gms of oxygen.

Example 5.9 Calculate the number of moles of Carbon-dioxide produced by the complete combustion of 10 moles of ethane.

Solution : Chemical equation for the combustion of Ethane

$$2\,C_2H_6\,(g) + 7\,O_2\,(g) \longrightarrow 4\,CO_2\,(g) + 6\,H_2O(g)$$

2 moles of ethane will give 4 moles of CO_2

1 mole of ethane will give $\frac{4}{2} = 2$ moles of CO_2

10 moles of ethane will give 2 × 10 = 20 moles of CO_2 **Ans.**

Example 5.10 Potassuim permanganate reacts with hydrochloric acid according to the equation.

$$2\,KMnO_4 + 16\,HCl \longrightarrow 2\,KCl + 2\,MnCl_2 + 4\,H_2O + 5\,Cl_2$$

Calculate the number of moles of chlorine gas that will be produced from 3·2 moles of hydrochloric acid.

Solution : According to the equation

$$2\,KMnO_4 + 16\,HCl \longrightarrow 2\,KCl + 2\,MnCl_2 + 4\,H_2O + 5\,Cl_2$$

16 moles of HCl will give 5 moles of Cl_2

1 mole of HCl will give $\frac{5}{16}$ moles of Cl_2

3·2 moles HCl will give $\frac{5}{16} \times 3{\cdot}2$ moles of HCl

= **1 mole of HCl Ans.**

Example 5.11 In electrolysis, one K^+ ion picks up one electron to give a K atom. Calculate the number of electrons required to obtain 39 gms of Potassuim. If the charge present on an electron is $1{\cdot}6 \times 10^{19}$ coulombs, calculate the total charge that 39 gms K^+ ions require during its conversion to K atoms. (Atomic mass of K = 39)

Solution : The reaction that takes place is

$K^+ + e \longrightarrow K$

1 mole of K^+ ions take up 1 mole of electrons to give 1 mole K atom

Atomic mass of K = 39

Gram atomic mass = 39 gm

So 39 gms of K contains 1 mole K atoms or $6{\cdot}023 \times 10^{23}$ atoms of K.

Hence 39 gms K atom is obtained by accepting $6{\cdot}023 \times 10^{23}$ electrons. **Ans.**

Charge on one electron = $1{\cdot}6 \times 10^{19}$ coulomb

So charge on $6{\cdot}023 \times 10^{23}$ electrons

= $1{\cdot}6 \times 10^{19} \times 6{\cdot}023 \times 10^{23}$ coulombs

= **$9{\cdot}6368 \times 10^{42}$ coulombs. Ans.**

Example 5.12 Atomic mass of silver is 107·87 a.m.u. Calculate the mass of a single atom of silver.

Solution :

Atomic mass of silver = 107·87 a.m.u.

So gram atomic mass of silver = 107·87 gms

Hence 1 mole of silver atoms weighs= 107·87 gms

$6{\cdot}23 \times 10^{23}$ atoms of silver weigh = 107·87 gms

1 atom of silver weighs = $\frac{107{\cdot}87}{6{\cdot}023 \times 10^{23}}$ gms.

= $17{\cdot}92 \times 10^{23}$ gms.

Example 5.13 The mass of a single atom of an element 'X' is $1{\cdot}7 \times 10^{-24}$ grams. Calculate its atomic weight ?

Solution :

Mass of one atom = $1{\cdot}7 \times 10^{-24}$ gms

Mass of $6{\cdot}023 \times 10^{23}$ atoms

= $1{\cdot}7 \times 10^{-24} \times 6{\cdot}023 \times 10^{23}$ gms

= 1·02371 gms

So mass of 1 mole atoms

= 1·02371 gms.

Hence gram atomic mass

= 1·02371 gm/mole

So atomic mass = **1·02371 a.m.u. Ans.**

SUMMARY

5.1 The force of attraction by virtue of which a substance enters into chemical combination with another is called *chemical affinity*.

5.2 All the chemical reactions can be classified under the following heads :

(A) *Decomposition reactions* such as Thermal decomposition, electrolysis, photochemical

(B) *Rearrangement*. Isomerisation

(C) Combination reactions

(D) Single displacement or Substitution reactions.

(E) Double decomposition reactions.

5.3 *Redox reactions* in which oxidation and reduction occur simultaneously.

5.4 Oxidation involves loss of electrons whereas reduction involve gain of electrons.

5.5 *Atomic mass* or *Atomic weight* : Atomic mass of an element is the relative mass of an atom of an element as compared with an atom of carbon^{-12} as the standard.

5.6 *Atomic mass unit* : One atomic mass unit is equal to one-twelfth of the mass of one carbon^{-12} atom.

5.7 *Molecular mass* or *Molecular weight* : The molecular mass is the relative mass of a molecule of a substance as compared to the one-twelfth of the atomic mass of one atom of carbon^{-12}.

5.8 *Gram atomic mass* or *Gram molecular mass* : Gram atomic mass is the atomic mass expressed in grams where gram molecular mass is the molecular mass in grams.

5.9 *Mole :* Mole is a chemists unit to represent a set of Avogadro's Number of particles Avogadro's number is a collection of $6{\cdot}023 \times 10^{23}$ particles.

5.10 Mole gives us a relationship between number and amount.

5.11 *Empirical formula :* The simplest formula that indicates the ratio of the numbers of different kinds of atoms present per molecule is known as *empirical formula.*

5.12 *Molecular formula :* The formula that indicates the actual number of each kind of atoms present per molecule is known as molecular formula.

5.13 $n = \dfrac{\text{Relative molecular mass}}{\text{Empirical formula mass}}$

5.14 *Chemical equation :* A chemical equation is the representation of a chemical change by means of symbols and formulas.

QUESTIONS

5.1 Given the following equations, label each as a combination : decomposition ; single displacement and double displacement reactions :

(*a*) $F_2 + 2\,KI \longrightarrow 2\,KF + I_2$

(*b*) $4\,Fe + 3O_2 \longrightarrow 2\,Fe_2O_3$

(*c*) $Zn + FeSO_4 \longrightarrow ZnSO_4 + Fe$

(*d*) $2\,H_2O_2 \longrightarrow 2\,H_2O + O_2$

(*e*) $N_2O_4 \longrightarrow 2\,NO_2$

(*f*) $C_2H_4 + H_2 \longrightarrow C_2H_6$

(*g*) $CuSO_4 + BaCl_2 \longrightarrow CuCl_2 + BaSO_4$

(*h*) $2\,AgBr \xrightarrow{\text{light}} 2\,Ag + Br_2$

(*i*) $NH_4NO_2 \longrightarrow N_2 + 2H_2O$

5.2 Which of the following reactions are redox reactions ? Identify the reducing agent and the oxidising agent in these redox reactions :

(*a*) $Pb^{2+}(aq) + 2I^-(aq) \longrightarrow PbI_2(s)$

(*b*) $NH_3(g) + HCl(g) \longrightarrow NH_4Cl(s)$

(*c*) $CH_4(g) + 2\,Cl_2(g) \longrightarrow C(s) + 4\,HCl(g)$

(*d*) $2\,SO_2(g) + O_2(g) \longrightarrow 2\,SO_3(g)$

(*e*) $Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$

(*f*) $Cu(s) + 2\,AgNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2\,Ag(s)$

(*g*) $2\,KI(aq) + Br_2(g) \longrightarrow 2\,KBr(aq) + I_2(s)$

(*h*) $4\,Fe(s) + 3O_2(g) \longrightarrow 2\,Fe_2O_3(s)$

(*i*) $CH_4(g) + 2F_2(g) \longrightarrow C(s) + 4\,HF(l)$

5.3 Choose the true statements in the following :

(*a*) In the reaction, $4\,Fe + 3O_2 \longrightarrow 2\,Fe_2O_3$, oxygen is the oxidising agent and iron gets reduced.

(*b*) A substance that can supply electrons to other substances is a reducing agent.

(*c*) The reaction $F_2 + 2\,NaI \longrightarrow 2\,NaF + I_2$ is a double displacement reaction.

(*d*) Gain of electrons by an atom of an element is known as oxidation.

(*e*) A chemical equation represents the actual chemical change.

(*f*) Mole is a collection of $6 \cdot 023 \times 10^{23}$ particles.

(*g*) In oxidation reduction reaction, the number of electrons lost or gained must be equal.

(*h*) One atomic mass unit is equal to one-twelfth of the mass of an atom of carbon−12.

(*i*) The force of attraction by virtue of which a substance enters into chemical combination with another is known as chemical affinity.

(*j*) Empirical formula of C_6H_6 is C_2H_2.

(*k*) The empirical formula and molecular formula of covalent compounds is the same.

(*l*) A chemical reaction transforms one or more substances into a different set of substances.

5.4 Fill in the blanks with suitable words :

(*a*) Decomposition of silver bromide in the presence of light is an example of .... decomposition reaction.

(*b*) Potassium chlorate on heating decomposes to give potassium chloride and oxygen. This is an example of .... decomposition reaction.

(*c*) Burning of Magnesium in the presence of air is an example .... reaction.

(*d*) The temporary changes are known as .... changes.

(*e*) Chemical reactions can be represented with the help of .... and ....

(*f*) .... is the force behind all the chemical reactions.

(*g*) The reactions that involve the change in the internal structure of a substance resulting in the formation of a new substance are known as .... reactions.

(*h*) Compounds having the same molecular formula but different structural arrangements are called ....

(*i*) A combination reaction in which many small molecules combine together to form a single large molecule is known as ....

(*j*) The activity of the metallic elements .... down the group.

(*k*) An oxidising agent is one that .... electrons whereas a reducing agent is one that .... electrons.

(*l*) An oxidising agent gets .... and a reducing agent gets ....

(*m*) A substance is reduced whenever it .... oxygen.

(*n*) One atomic mass unit is equal to one-twelfth of the mass of one .... atom.

(*o*) The sum of the atomic masses of the atoms in the formula of a molecule is its ....

(*p*) The formula for all electrovalent compounds are .... formulae.

5.5 Match the terms in Column A with the descriptions in Column B.

Column A	Column B
1. Chemical reaction	1. starting substances.
2. Chemical equation	2. substances produced.
3. Reactant	3. donates electrons.
4. Product	4. accepts electrons.
5. Oxidising agent	5. one-twelfth the mass of an atom of C^{-12}.
6. Reducing agent	6. energy is either evolved or absorbed.
7. Atomic mass unit	7. short hand representative of chemical reaction.

5.6 Translate the following statements into chemical equations and also balance these equations :

1. Hydrogen gas burns in oxygen gas to give water.

2. Barium chloride reacts with sodium sulphate to give a precipitate of Barium sulphate and sodium chloride.
3. Hydrogen sulphide gas burns in the presence of limited air to give water and sulphur.
4. Aluminium metal replaces iron from ferric oxide to give aluminium oxide and iron.
5. Potassium nitrate on heating decomposes to give potassium nitrite and oxygen.
6. On passing electricity through acidulated water hydrogen and oxygen is produced.
7. Potassium chlorate on heating decomposes to give potassium chloride and oxygen.
8. When a strip of Zinc metal is dipped in a solution of copper sulphate, copper is precipitated and a solution of Zinc sulphate is obtained.
9. Aluminium sulphate solution when mixed with ammonium hydroxide solution gives gelatinous white precipitate of aluminium hydroxide and ammonium sulphate solution.
10. On adding sodium hydroxide solution to an aqueous solution of copper sulphate, sodium sulphate and a precipitate of copper hydroxide is obtained.
11. Ammonia burns in oxygen to give nitrogen and water.
12. Copper when heated with conc. nitric acid gives copper nitrate, nitrogen dioxide and oxygen gas.
13. Hydrochloric acid when added to an aqueous solution of sodium carbonate gives an aqueous solution of sodium chloride, carbon dioxide gas and water.
14. Magnesium burns in a jar of carbon dioxide to give magnesium oxide and carbon flakes.
15. On adding sulphuric acid to aluminium oxide solid, an aqueous solution of aluminium sulphate and water is obtained.

5.7 Balance the following chemical equations :

1. $PtCl_4 \longrightarrow Pt + Cl_2$
2. $P_4 + O_2 \longrightarrow P_4O_{10}$
3. $KNO_3 \longrightarrow KNO_2 + O_2$
4. $Cu(NO_3)_2 \longrightarrow CuO + NO_2 + O_2$
5. $H_3PO_3 \longrightarrow H_3PO_4 + PH_3$
6. $Pb + H_2O + O_2 \longrightarrow Pb(OH)_2$
7. $PCl_5 + H_2O \longrightarrow POCl_3 + HCl$
8. $CaO + C \longrightarrow CaC_2 + CO$
9. $Al_2(SO_4)_3 + NH_4OH \longrightarrow Al(OH)_3 + (NH_4)_2SO_4$
10. $CuSO_4 + NaOH \longrightarrow Cu(OH)_2 + Na_2SO_4$
11. $KMnO_4 + H_2SO_4 + FeSO_4 \longrightarrow K_2SO_4 + Fe_2(SO_4)_3 + H_2O$
12. $MnO_2 + HCl \longrightarrow MnCl_2 + H_2O$
13. $HCl + O_2 \longrightarrow H_2O + Cl_2$
14. $KMnO_4 + HCl \longrightarrow KCl + MnCl_3 + H_2O + Cl_2$
15. $CaCl_2 + K_3PO_4 \longrightarrow Ca_3(PO_4)_2 + KCl$
16. $C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$
17. $Fe + Cl_2 \longrightarrow FeCl_3$
18. $Ca(OH)_2 + NH_4Cl \longrightarrow CaCl_2 + NH_3 + H_2O$
19. $Na + H_2O \longrightarrow NaOH + H_2$

20. $S + H_2SO_4 \longrightarrow SO_2 + H_2O$.

5.8 In what way the two reactions in each of the following pairs are different from each other ?

(*a*) (*i*) $Na_2O\,(s) + H_2O\,(l) \longrightarrow 2\,NaOH\,(aq)$

(*ii*) $2Mg\,(s) + O_2\,(g) \longrightarrow 2\,MgO\,(s)$

(*b*) (*i*) $Zn\,(s) + FeSO_4\,(aq) \longrightarrow ZnSO_4\,(aq) + Fe\,(s)$

(*ii*) $H_2S\,(aq) + CuSO_4\,(aq) \longrightarrow CuS(s) + H_2SO_4\,(aq)$

(*c*) (*i*) $MgCO_3(s) \xrightarrow{heat} MgO(s) + CO_2\,(g)$

(*ii*) $2AgBr \xrightarrow{light} 2\,Ag + Br_2$

(*d*) (*i*) $2H_2O\,(l) \xrightarrow[Current]{Electric} 2H_2(g) + O_2\,(g)$

(*ii*) $NH_4Cl\,(s) \xrightarrow{heat} NH_3(g) + HCl\,(g)$

(*e*) (*i*) $NaOH\,(aq) + HCl\,(aq) \rightarrow NaCl\,(aq) + H_2O(l)$

(*ii*) $AgNO_3\,(aq) + NaCl\,(aq) \rightarrow AgCl\,(s) + NaNO_3(aq)$

(*f*) (*i*) $BaCl_2\,(aq) + H_2SO_4\,(aq) \rightarrow BaSO_4\,(s) + 2HCl\,(aq)$

(*ii*) $FeCl_3\,(aq) + 3HOH\,(l) \rightarrow Fe(OH)_3\,(s) + 3HCl\,(aq)$.

5.9 Classify the following reactions into combination, decomposition ; single displacement and double displacement reactions :

(*a*) $NaCl\,(aq) + AgNO_3\,(aq) \rightarrow AgCl\,(s) + NaNO_3\,(aq)$

(*b*) $Zn\,(s) + CuSO_4\,(aq) \rightarrow ZnSO_4\,(aq) + Cu\,(S)$

(*c*) $2\,CO\,(g) + O_2\,(g) \rightarrow 2\,CO_2\,(g)$

(*d*) $NH_4Cl\,(s) \rightarrow NH_3\,(g) + HCl\,(g)$

(*e*) $2\,H_2\,(g) + O_2(g) \rightarrow 2\,H_2O\,(l)$

(*f*) $H_2S\,(g) + Cl_2\,(aq) \rightarrow 2\,HCl\,(aq) + S(s)$

(*g*) $CaCO_{-}\,(s) \rightarrow CaO(s) + CO_2\,(g)$

(*h*) $NH_4NO_2\,(s) \rightarrow N_2\,(g) + 2H_2O\,(l)$

(*i*) $2\,Al\,(s) + Fe_2O_3\,(s) \rightarrow Al_2O_3\,(s) + 2\,Fe\,(l)$

(*j*) $PbCl_2\,(aq) + H_2SO_4\,(aq) \rightarrow PbSO_4\,(s) + 2\,HCl\,(aq)$

(*k*) $2\,NH_3\,(g) + H_2SO_4\,(aq) \rightarrow (NH_4)_2SO_4\,(aq)$.

5.10 What are redox reactions ? Identify the redox reactions from the followings :

(*a*) $Zn\,(s) + Cl_2\,(g) \rightarrow ZnCl_2\,(s)$

(*b*) $CuSO_4\,(aq) + Zn\,(s) \rightarrow ZnSO_4\,(aq) + Cu\,(s)$

(*c*) $NH_3\,(g) + HCl\,(g) \rightarrow NH_4Cl\,(s)$

(*d*) $Zn\,(s) + Fe^{++}\,(aq) \rightarrow Zn^{++}\,(aq) + Fe\,(s)$

(*e*) $2H_2O\,(l) \rightarrow 2H_2\,(g) + O_2\,(g)$

(*f*) $CuSO_4\,(aq) + H_2S\,(aq) \rightarrow CuS\,(s) + H_2SO_4\,(aq)$

(*g*) $2Mg\,(s) + O_2(g) \rightarrow 2MgO(s)$

5.11 From the following reactions identify the ordering and reducing agents.

(*a*) $Zn\,(s) + Fe^{++}\,(aq) \rightarrow Zn^{++}\,(aq) + Fe\,(s)$

(*b*) $CuO (s) + H_2 (g) \rightarrow Cu (s) + H_2O (q)$
(*c*) $2Fe (s) + 3 Cl_2 (g) \rightarrow 2 FCl_3 (s)$
(*d*) $Ca (s) + Cl_2 (g) \rightarrow Ca^{2+} (Cl^-)^2 (s)$
(*e*) $2 Al (l) + Fe_2O_3 (s) \rightarrow Al_2O_3 (s) + 2Fe (s)$

5.12 Write short notes on :
(*a*) Decomposition reaction
(*b*) Combination reaction
(*c*) Redox reaction
(*d*) Gram molecular mass
(*e*) Atomic mass unit

5.13 List the various types of double decomposition reaction. Give two examples of each type

5.14 Justify the statement "Oxidation and reduction occur simultaneously"

5.15 Explain the following :
(*a*) Ebcholysis is a type of redox reaction
(*b*) Blue colour of copper sulphate solution fades away when zinc rod is dipped in it.
(*c*) Burning is an oxidation reaction
(*d*) 2 gram of hydrogen contain the same number of molecules as 18 gram of water.

5.16 What is a chemical equation ? List the limitations of a chemical equation.

PROBLEMS

Note : For all the questions consult Table 5.1 for relative atomic masses

5.1 Calculate the formula weight of each of the following :
N_2 ; H_3PO_3 ; $Ca(NO_3)_2$; $C_6H_{12}O_6$; $Ba_3(PO_4)_2$;
NH_4 ; $(NH_4)_2SO_4$; $Al(OH)_3$; $ZnCl_2$; $KClO_3$.

5.2 Calculate the formula weight of each of the following minerals :
(*a*) Silver glance, Ag_2S
(*b*) Cryolite, Na_3AlF_6
(*c*) Lime stone, $CaCO_3$
(*d*) Copper pyrites, $CuFeS_2$
(*e*) Chromite, $FeO\ Cr_2O_3$
(*f*) Carnallite, $KCl\ MgCl_2 . 6\ H_2O$
(*g*) Dolomite $CaCO_3$, $MgCO_3$.
(*h*) Malachite, $Cu(OH)_2\ CuCO_3$.
(*i*) Chile salt petre, $NaNO_3$.

5.3 Calculate the weight of :
(*a*) 0·40 mole of Carbon dioxide
(*b*) 3 mole of K_3PO_4
(*c*) 0·04 mole of Chlorine gas (Cl_2)
(*d*) 4 gram atoms of chlorine
(*e*) 0·45 mole of $(NH_4)_2CO_3$
(*f*) 3·14 mole of H_5IO_6
(*g*) 2·0 mole of NH_3
(*h*) 2·53 moles of water
(*i*) 2 gram atom of sodium.

5.4 Calculate the number of moles in :
(*a*) $6{\cdot}023 \times 10^{10}$ molecules of water.
(*b*) $60{\cdot}23 \times 10^{9}$ molecules of sulphuric acid.
(*c*) $3{\cdot}0115 \times 10^{20}$ atoms of sodium.
(*d*) $1{\cdot}50575 \times 10^{23}$ atoms of oxygen.

5.5 Calculate the number of moles in :
(*a*) 3·4 gms of H_2O_2
(*b*) 2·8 gms of N_2
(*c*) 1·01 gms of KNO_3
(*d*) 11·6 gms of $(NH_4)_2SO_4$
(*e*) 6·8 gms of NH_3.

5.6 How many molecules of ammonia are present in 0·060 moles of ammonia ?

5.7 Calculate the weight of one atom of :
(*a*) Lead
(*b*) Nitrogen
(*c*) Silver
(*d*) Hg
from their gram atomic weights and Avogadro's Number.

5.8 Calculate each of the following :

(*a*) The percentage of Nitrogen in $NH_4\,NO_3$

(*b*) The percentage of Nitrogen, of Hydrogen ; of Sulphur and of Oxygen in $(NH_4)_2\,SO_4$

(*c*) The percentage of water in $ZnSO_4\;7H_2O$.

(*d*) The percentage of Ammonia in $COCl_3.\;6\,NH_3$.

5.9 Calculate the percentage composition of each of the following :

(*a*) $Al\,(OH)_3$

(*b*) $(CH_3COO)_2\,Pb$

(*c*) HgO

(*d*) $K_2\,Cr_2\,O_7$

(*e*) Nitric acid

5.10 Determine the emperical formula of a compound which has the following percentage composition. Carbon – 24.3% ; Hydrogen 4.1%; Chlorine 71·6% If molecular weight of the compound is 99, what is its molecular formula ?

5.11 What is the emperical formule of the compounds whose percentage composition is as follows :

(*a*) C = 41·35 ; H = 6·89 ; N = 24·12 ; O = 27·64

(*b*) C = 50·4 ; H = 9·25 ; O= 40·35

(*c*) C = 60 ; H = 13·3 ; O = 26·7

(*d*) C = 92·32 ; H = 7·68

5.12 In a balanced equation

$CH_4(g) + 2O_2\,(g) \rightarrow CO_2\,(g) + 2H_2O\,(l)$

(*a*) How many moles of methane will be required to obtain 8 moles of water ?

(*b*) How many grams of CO_2 will be obtained if you start the reaction with 4 moles of methane ?

5.13 In a balanced chemical equation :

$2KClO_3 \rightarrow 2KCl + 3O_2$

(*a*) How many moles of potassium chlorate will be required to obtain 6 moles of oxygen ?

(*b*) How many moles of oxygen gas will be evolved from 12·25 gms of $KClO_3$?

5.14 In the balanced chemical reaction

$CS_2(l) + 3O_2\,(g) \rightarrow CO_2\,(g) + 2SO_2\,(g)$

(*a*) How many moles oxygen are required for complete combustion of 0·025 moles of CS_2 ?

(*b*) How many moles of SO_2 gas are produced during the complete combustion of 0·025 moles of CS_2 ?

5.15 In the reaction

$2\,Na\,(s) + 2\,H_2O\,(l) \rightarrow 2Na\,OH\,(aq) + H_2\,(g)$

(*a*) How many grams of NaOH is obtained from 2·3 grams of sodium ?

(*b*) How many moles of hydrogen gas is evolved if 20 gms of NaOH is obtained ?

5.16 (A) Find out the ratio in which molecules are present in each of the following sets of substances.

(*a*) 12 litres of oxygen and 6 litres of chlorine at S.T.P

(*b*) 9 gms of water and 4·4 gms of CO_2

(*c*) 3·55 gm of chlorine gas and 3·2 gms of oxygen.

5.17 How many moles of Chlorine atoms are present in one mole of the following compounds :

(*a*) $Cl_2\,O_7$

(*b*) HCl

(*c*) $FeCl_3$

(*d*) $BaCl_2$

(*e*) $CHCl_3$.

□□

CHAPTER 6

ENERGY CHANGES IN CHEMICAL REACTIONS

6.1 INTRODUCTION :

Energy is the most precious commodity that we have. It gives us the power to do work. There can be no life, no warmth and no movement without energy. Living standard of any country's people is related to the availability of energy resources. Do you know man has been worshipping the Sun from the earliest time? This is because Sun provides us with the most of our food through the process of photosynthesis. It has created our supply of fossil fuels such as coal, oil and natural gas. The transfer of energy to and from the chemicals play an important role in the chemical processes in industry and in living beings. Hence the study of energy changes are important for us. During the burning of a fuel, chemical energy is converted into heat energy. Even within our body energy changes are very important. Carbohydrates and fats are an important biological fuels. As a result of the metabolism going on in our body the chemical energy of the food is being converted to heat energy so as to keep the body warm; to electrical energy to carry out the signals in our nerve fibres; to mechanical energy in our muscles.

In our day-to-day life we observe that chemical energy is being converted to electrical energy in the cells and batteries; to heat and light energy in the chemical reactions and so on. In the present unit we shall discuss about all. What is the driving force of these energies ? What is it that causes a reaction to occur ?

6.2 EXOTHERMIC AND ENDOTHERMIC REACTIONS

Every substance has a definite amount of energy associated with it. The amount of energy associated with any substance is different for different substances. During a chemical reaction rearrangement of atoms take place. So the energy associated with the substance before the reaction is different from the energy after the reaction. This is because the energy of the reactant is always different from that of the products. These chemical changes are generally shown by the change in temperature. If during a chemical reaction heat is given out then the reactions are known as *Exothermic Reactions.* In the case of exothermic reactions temperature of the products rises. In such cases the energy of the products is less than the reactants and the difference of energies of the products and reactants is given out in the form of heat.

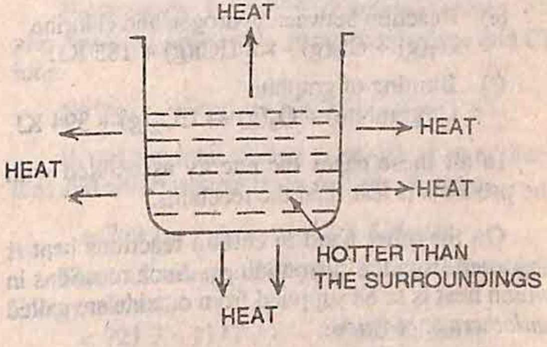


Fig. 6·1 In an exothermic reaction heat is lost to the surroundings.

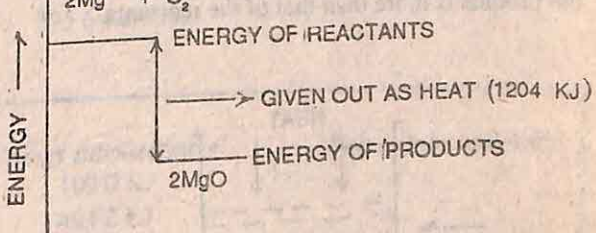


Fig. 6.2

For Example : When magnesium reacts with oxygen, heat is given out.

$$2\,Mg(s) + O_2(g) \rightarrow 2MgO(s) + 1204\ KJ.$$

A part of the chemical energy in magnesium and oxygen is changed into the chemical energy in MgO and the remaining is released in the form of heat. Hence we can say chemical energy in magnesium oxide is less than magnesium and oxygen.

A few more examples are :

(*a*) The reaction between hydrogen and oxygen to produce water and heat. Hydrogen and oxygen have higher total energy content than water.
$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l) + \text{heat.}$

(*b*) Burning of methane
$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) + 890{\cdot}4\ KJ.$

(*c*) Dissociation of hydrogen iodide
$2\ HI(g) \rightarrow H_2(g) + I_2(s) + 52{\cdot}5\ KJ.$

(*d*) Reaction between acid and base
$H^+(aq) + OH^-(aq) \rightarrow H_2O(l) + 57.1\ KJ.$

(*e*) Reaction between hydrogen and chlorine
$H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g) + 185\ KJ.$

(*f*) Burning of graphite
$C\ (\text{graphite}) + O_2(g) \rightarrow CO_2(g) + 394\ KJ.$

In all these cases the energy associated with the products is less than the reactants.

On the other hand in certain reactions heat is absorbed from the surroundings. Such reactions in which heat is to be supplied from outside are called *endothermic reactions.*

In the case of endothermic reactions final temperature of the products is below the initial temperature. In these total energy associated with the produts is more than that of the reactants.

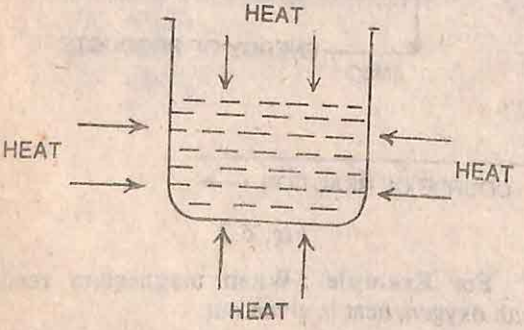


Fig. 6.3 *In an endothermic reaction heat is absorbed from the surroundings.*

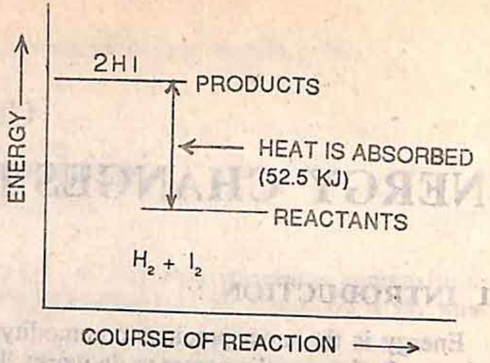


Fig. 6.4

During an endothermic reaction Reactants having low energy content change into products having high energy content.

For example : Combination of hydrogen and iodine is carried out with the absorption of heat.

$H_2(g) + I_2\ (s) \rightarrow 2\ HI(g) + 52.5\ K.J.$

The total energy content of HI is more than that of hydrogen and iodine.

Hence any reaction in which heat is taken in from the surrounding by the system are known as endothermic reactions.

In such cases temperature of the surroundings is lowered.

A few more examples are:

(A) Combination of Nitrogen gas with oxygen gas to produce nitric oxide.
$N_2(g) + O_2(g) \rightarrow 2\ NO(g) - 180{\cdot}5\ KJ$

(B) Decomposition of Ammonia
$2\ NH_3(g) \rightarrow N_2(g) + 3H_2(g) - 92.8\ KJ$

(C) Production of water gas.
$C(s) + H_2O(g) \rightarrow CO(g) + H_2(g) - \text{heat}$

In a chemical reaction old bonds are broken and new bonds are formed. Energy is needed to break the bond and energy is released when bonds are formed. Each bond has certain definite amount of energy associated with it. The difference of these energies is either given out or absorbed. A reaction is exothermic if the energy liberated in the bond formation process is more than that required for the breaking of bonds. And it is endothermic if the energy liberated in the bond formation process is less than that required for the breaking up of bonds.

Units of Heat : Heat is measured in calories and kilo calories. One calorie is the amount of heat required to raise the temperature of one gram of water by 1 degree celsius temperature.

1 K.Cal = 1000 cal.

The unit of energy now-a-days is Joule and kilojoule abbreviated as J and KJ respectively.

1 calorie = 4·2 Joules.

6.3 BOND ENERGY

We know when two atoms approach near each other to form a compound energy is liberated. The amount of energy that is liberated when two atoms combine to form a molecule is known as *Bond Energy*. The same amount of energy has to be supplied from outside to break up the bond. The amount of energy required to break up the bond is known as *Bond dissociation energy*. Hence we can say that each type of bond is associated with a definite amount of energy known as Bond Energy. This energy is absorbed when a bond is broken and evolved when the bond is formed. So energy change in a chemical reaction is the difference of the energies between the bond making and bond breaking. Bond energies can be added and subtracted just like algebraic addition and subtraction.

Let us calculate the energy changes taking place during the formation of sodium chloride from its elements, sodium and chlorine.

The reaction is $2\,Na(s) + Cl_2(g) \rightarrow 2\,NaCl(s)$

or $Na(s) + \frac{1}{2}Cl_2(g) \rightarrow Na\,Cl(s)$

In this case the energy is to be supplied for converting solid sodium metal to sodium atoms and sodium atoms are then converted to sodium ions (Na^+)

$$Na(s) + 109\ KJ \rightarrow Na(g)$$

$$Na(g) + 493\ 2\ KJ \rightarrow Na^+(g) + e^-$$

So the total energy supplied for converting Na (s) to Na^+ (g)

$$= 493{\cdot}2\ KJ + 109\ KJ = 602.2\ KJ.$$

Energy is also supplied for converting chlorine gas to chloride ions (Cl^-)

$$Cl_2(g) + 239.0\ KJ = 2\,Cl^-(g)$$

$$\frac{1}{2}Cl_2(g) + 119{\cdot}5\ KJ \rightarrow Cl^-(g)$$

So the total energy supplied during the reaction is

= 602·2 KJ/mole + 119·5 KJ/mole

= 721·7 KJ/mole.

But energy is released when chlorine atoms pick an electron (Electron Affinity)

$$Cl(g) + \bar{e} \rightarrow Cl^-(g) + 348{\cdot}2\ K.J.$$

Electrostatic forces of attractions comes into play between oppositely charged Na^+ ions and Cl^- ions.

$$Na^+(g) + Cl^-(g) \rightarrow NaCl(s) + 769.1\ KJ.$$

Hence the total energy released when sodium ions and chlorine atom form sodium chloride

= 769·1 + 348·2 = 1117·3 KJ/mole

The total energy change = Energy supplied - Energy released

= 721·7 – 1117·3 = — 395·6 KJ/mole.

Hence in this case 395·6 KJ/mole of energy is given out.

So Energy of formation of sodium chloride is 395·6 KJ/mole.

It can be summed up as follows :

Reaction	*Energy supplied/mole*	*Energy released/mole*
(*i*) $Na(s) \rightarrow Na(g)$	109·0 KJ	
(*ii*) $Na(g) \rightarrow Na^+(g) + \bar{e}$	493·2 KJ	
(*iii*) $\frac{1}{2}Cl_2(g) \rightarrow Cl(g)$	119·5 KJ	
Total	721·7 KJ	
(*iv*) $Cl(g) + \bar{e} \rightarrow Cl^-(g)$		348·2 KJ
(*v*) $Na^+(g)\,Cl^-(g) \rightarrow NaCl(s)$		769·1 KJ
		Total 1117·3 KJ

Energy change 721.7 KJ – 1117.3KJ = – 395.6 KJ/mole

Example 6·1 :

Q. What do you understand by the term "Energy of formation of sodium chloride is 395·6 KJ/mole".

Ans : It means that 395·6 KJ. of energy is released when 1 mole of sodium chloride is formed from its elements *i.e.* , sodium and chlorine gas.

Problem:

Calculate the energy of formation of water from hydrogen and oxygen.

Solution :

The different reactions that take place are :

$$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(g)$$

So the first step is

$$H_2(g) \rightarrow 2\,H(g)$$

Energy is to be supplied for it

$$H_2(g) + 435{\cdot}6 \text{ KJ/mole} \rightarrow 2\,H(g) \qquad -(i)$$

Second step is $O_2(g) \rightarrow 2O(g)$

$$O_2(g) + 494{\cdot}6 \text{ KJ/mole atom} \rightarrow 2O(g)$$

$$\text{or} \quad \frac{1}{2}O_2(g) + 247{\cdot}3 \text{ KJ/mole atom} \rightarrow O(g) \quad -(ii)$$

Energy will be released when H and O combine to give $H_2O(g)$

$$2H(g) + O(g) \rightarrow 2H_2O(g) + 924{\cdot}4 \text{ KJ/mole.} \; (iii)$$

Adding (*i*), (*ii*) and (*iii*) on the L.H.S & R.H.S.

$$H_2(g) + 435{\cdot}6 \text{ KJ/mole} + \frac{1}{2}O_2(g) + 247{\cdot}3 \text{ KJ/mole} + 2H + O \rightarrow 2H(g) + O(g) + H_2O(g) + 924{\cdot}4 \text{ KJ/mole}$$

$$\text{or} \quad H_2(g) + \frac{1}{2}O_2(g) + 682{\cdot}9 \text{ KJ/mole} \rightarrow H_2O(g) + 924{\cdot}4 \text{ KJ/mole}$$

$$\text{or} \quad H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(g) + (924{\cdot}4 - 682{\cdot}9) \text{ KJ/mole}$$

$$\text{or} \quad H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(g) + 241{\cdot}5 \text{ KJ/mole}$$

Ans.

So the energy of formation of water is 241·5 KJ/mole.

It means 241·5 KJ/mole of energy is released when water is formed from its elements $H_2(g)$ and $O_2(g)$.

$$\text{Hence} \left.\begin{array}{l}\text{Energy}\\ \text{change in}\\ \text{a chemical}\\ \text{reaction}\end{array}\right\} = \left\{\begin{array}{l}\text{Energy}\\ \text{required}\\ \text{to break}\\ \text{the bonds}\\ \text{in the}\\ \text{molecules}\\ \text{of the}\\ \text{reactants}\end{array}\right\} \text{minus} \left\{\begin{array}{l}\text{Engery}\\ \text{released}\\ \text{when bonds}\\ \text{in the mole-}\\ \text{cules of the}\\ \text{products}\\ \text{are formed.}\end{array}\right.$$

Examples 6.2 :

Energy of breaking up of H_2 bond and Cl_2 bonds are 437 KJ/mole and 244 KJ/mole respectively. And energy of formation of HCl is 433 KJ/mole. Calculate the energy of the reaction $H_2 + Cl_2 \rightarrow 2\,HCl$.

Solution :

Step I. Energy for breaking up of H_2 bond is 437 KJ/mole

$$\text{Hence } H_2(g) + 437 \text{ KJ/mole} \rightarrow 2H(g) \quad \ldots (i)$$

Step II Energy for breaking up of Cl_2 bond is 244 KJ/mole.

$$\text{Hence } Cl_2(g) + 244 \text{ KJ/mole} \rightarrow 2\,Cl(g) .. (ii)$$

Step III Energy of bond formation of HCl(g) is 433 KJ/mole

$$H(g) + Cl(g) \rightarrow HCl\,(g) + 433 \text{ KJ/mole}$$

$$\text{or } 2\,H(g) + 2\,Cl(g) \rightarrow 2\,HCl(g) + 866 \text{ KJ} \ldots (iii)$$

(This is in order to obtain 2 HCl in the products)

Now add (*i*), (*ii*) and (*iii*) to obtain the required equation $H_2 + Cl_2 \rightarrow 2HCl$

$$H_2(g) + Cl_2(g) + (437 + 244 \text{ KJ}) + 2H(g) + 2Cl(g) \rightarrow 2H(g) + 2\,Cl(g) + 2HCl(g) + 866 \text{ KJ.}$$

$$\text{or } H_2(g) + Cl_2(g) + 681 \text{ KJ} \rightarrow 2\,HCl(g) + 866 \text{ KJ}$$

$$\text{or } H_2(g) + Cl_2(g) \rightarrow 2\,HCl(g) + (866 - 681) \text{ KJ}$$

$$\text{or } H_2(g) + Cl_2(g) \rightarrow 2\,HCl(g) + 185 \text{ KJ}$$

Example 6·3 :

For the reaction

$$C(s) + O_2(g) \rightarrow CO_2(g) + 395 \text{ KJ/mole}$$

Calculate the heat evolved when $\frac{1}{2}$ Kg of carbon is burnt (Atomic masses of C = 12).

Solution :

$$C(s) + O_2(g) \rightarrow CO_2(g) + 395 \text{ KJ/mole}$$

So 1 mole of carbon gives 395 KJ heat

1 mole carbon = 12 gm.

Hence 12 gms. of carbon gives 395 KJ heat

1 gm. of carbon gives $\frac{395}{12}$ KJ heat

$\frac{1}{2}$ Kg or 500 gm. of carbon gives $\frac{395}{12} \times 500$ KJ heat

$= \frac{197500}{12}$ KJ heat

$= 16458{\cdot}33$ KJ heat

Example 6·4 :

Methane burns in air to give carbon dioxide and water according to the equation

$$CH_4(g)+2O_2(g) \rightarrow CO_2(g)+2H_2O(l) + 890 \text{ KJ/mole}$$

Calculate the amount of heat energy released when 80 gms. of methane is burnt.

Solution :

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2\,H_2O(l) + 890$ KJ/mole

1 mole of $CH_4 = (12 + 1 \times 4)$ gm. = 16 gms.

1 mole of methane on burning gives 890 KJ heat

or 16 gm of methane on burning gives 890 KJ heat

1 gm. of methane on burning gives $\frac{890}{16}$ KJ heat

80 gm. of methane on burning will give

$\frac{890}{\cancel{16}} \times \cancel{80}^{5}$ KJ = **4450 KJ Ans.**

Example 6·5 :

The heat of combustion of butane is 2890 KJ/mole. What will be the amount of heat liberated when 1 Kg of butane is burnt in the presence of excess of air ?

Solution:

Butane burns in air according to the equation

$$C_4H_{10}(g) + 13/2\,O_2(g) \rightarrow 4CO_2(g) + 5H_2O(g) + 2890 \text{ KJ}$$

1 mole butane gives 2890 KJ heat

1 mole butane = $(4 \times 12 + 10 \times 1)$ gm = 58 gm.

Hence 58 gms. of butane gives 2890 KJ heat

1 gm. of butane gives $\frac{2890}{58}$ KJ heat

1000 gm. of butane will give $\frac{2890}{58} \times 1000$ KJ

= **49828 KJ. Ans.**

Example 6·6 :

The formation of NH_3 from its elements takes place according to the equation

$$N_2(g) + 3H_2(g) \rightarrow 2\,NH_3(g) + 92 \text{ KJ}$$

Calculate the heat of formation for 1 mole of NH_3.

Solution:

$$N_2(g) + 3H_2(g) \rightarrow 2\,NH_3(g) + 92 \text{ KJ}$$

1 mole of $N_2(g)$ combines with 3 mole of $H_2(g)$ to give 2 moles of $NH_3(g)$

Hence formation of 2 mole of $NH_3(g)$ gives 92 KJ heat

For 1 mole formation of $NH_3(g)$ heat given out will be $= \frac{92}{2}$ KJ = **46 KJ. Ans.**

So Heat of formation of NH_3 = 46 KJ/mole.

6.4 LIGHT ENERGY CHANGES TAKING PLACE IN A REACTION OR PHOTOCHEMICAL REACTIONS

Heat is not only the form of energy. Energy changes can also take place in the form of light. In a number of reactions light is either absorbed or evolved. Such reactions that are caused by the action of light are known as *Photochemical reactions*. In such reactions the light absorbed, provides the necessary energy to the molecules involved in the chemical reactions.

A few examples of this types are :

(*a*) It has been observed that colour of the cloth fades away when exposed to sunlight. This is because colour of the cloth is an organic compound and when exposed to sunlight, these molecules decompose.

(*b*) The reaction between H_2 and Cl_2 is very slow in the dark but if the mixture is exposed to sunlight, explosion takes place.

Cl_2 + light $\rightarrow$ 2 Cl
$Cl + H_2 \rightarrow HCl + H$
$H + Cl_2 \rightarrow HCl + Cl$ and so on.

Chlorine atoms are produced photochemically and the chain reaction starts.

(*c*) In the atmosphere also a number of reactions take place by the absorption of light. Gases like SO_2 and NO_2 that are given out as industrial pollutants absorb sunlight and produce free radicals. These free radicals react further with a number of molecules in the air to form a number of harmful compounds.

(*d*) In the atmosphere Ozone is formed by the absorption of ultraviolet light.

(*e*) ***Photography*** is the best example of a *photochemical reaction*. The photographic film contain silver bromide emulsion. When the photographic film is exposed to light, AgBr decomposes to Ag.

$$Ag^+ Br^- (s) \xrightarrow{\text{light}} Ag(s) + \frac{1}{2} Br_2(g) \rightarrow$$

During the development stage the film is treated with Hypo that removes the excess of AgBr as a soluble complex and Ag remains on the film as an opaque shadow.

(*f*) Freshly prepared precipitates of silver chloride when exposed to sunlight becomes purple grey.

(g) The reaction between halogen and alkane is also an example of a photochemical reaction.

Photosynthesis: Photosynthesis is an example of a photochemical reaction. Photosynthesis is the process of building up of carbohydrates in the green parts of the plants by chlorophyll in the presence of sunlight by using carbon dioxide and water. During photosynthesis light energy absorbed by the green parts of the plants is transferred to chemical energy stored up by the compounds. Oxygen is liberated as a bye-product.

$$6\,CO_2(g) + 6\,H_2O(l) \xrightarrow{\text{Sunlight}} C_6H_{12}O_6 + 6\,CO_2 - \text{heat.}$$

In the absence of sunlight photosynthesis ceases because energy is no longer provided.

Experiment : To show that oxygen is liberated during photosynthesis.

Take some green water plants in a beaker. Fill the beaker with water. Add small quantity of baking soda in water. Cover the plants with a funnel as shown in Fig. 6·5.

Invert a test tube filled with water over the funnel. Place the set up apparatus in sunlight. After sometime bubbles of a colourless gas are seen to rise upwards and collect in the test tube. Test the gas collected in the test tube and it is found to be oxygen.

During the process of respiration this chemical energy is converted into heat energy.

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6\,CO_2 + 6\,H_2O + \text{heat}$$

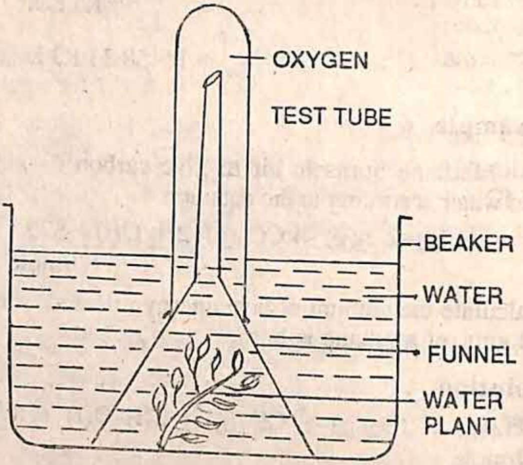


Fig. 6.5

Electrochemistry: Uptill now we have studied the conversion of chemical energy into heat energy and vice versa or chemcal energy into light energy and vice versa or chemical energy into mechanical energy and vice versa. Now we will study about the conversion of *chemical energy into electrical energy* and vice versa.

Electrochemistry is the study of chemical reactions in which electrical energy is converted into chemical energy and vice versa. All the electrochemical reactions can be classified under one of the following heads:

(A) Those reactions in which chemical reactions are carried out at the cost of electricity *i.e.*, electrical energy is converted into chemical energy. Cells used for this purpose are called *Electrolytic cells.*

(B) The other type of reactions are those in which electricity is produced at the cost of chemical reactions *i.e.*, chemical energy is converted into electrical energy. Cells of this type are known as *Electrochemical cells.*

All the electrochemical reactions are redox reactions *i.e.*, oxidation and reduction takes place simultaneously.

Before discussing about electrochemical

reactions let us study the following terms:

(a) **Electrolyte:** An electrolyte is a compound which when in solution or melted, conducts an electric current and is decomposed by it.
Electrolyte consists of positively charged and negatively charged ions.

(b) **Non-electrolyte** : It is a compound which, when in solution or melted, does not conduct an electric current *e.g.*, Sugar solution.

(c) **Electrolysis** : It is the decomposition of an electrolyte by passing an electric current through it.

(d) **Electrodes** : These are two pieces of metal or graphite by which electric current *i.e.*, electron enters and leaves an electrolyte.

(e) **Cathode:** It is the negative electrode by which electrons enter an electrode. *Anode* is the positive electrode by which electrons leave the electrode. *Anode for oxidation and Cathode for reduction is a rule in electrochemistry.*
units of Electric current:

Ampere : It is a unit of current. It is the unit for measuring the flow of electrons in a circuit. So it measures the rate of flow of electricity.

Coulomb : It is the quantity of electricity. The quantity of electricity, when a current of 1 ampere flows for 1 second, is known as one *coulomb*. So when I ampere current flows for t seconds then,

$$\begin{aligned}\text{Quantity of electricity} &= 1 \text{ ampere} \times t \text{ seconds}\\ &= I \times t \text{ amp. sec.}\\ &= I\,.\,t \text{ coulomb.}\end{aligned}$$

So coulombs = Ampere × Second.

Faraday : One faraday is the quantity of electricity required to produce 1 gm. atom of a monovalent atom OR it is the quantity of electricity required to liberate 1 gm. equivalent of an element from its solutions.

The number of *faradays* required to liberate one mole of any element tells about the charge on the ions of that element.

For example:

$$Ag^+ + e \rightarrow Ag$$

So the quantity of electricity required to liberate one mole of Ag from its solution will be one faraday.

Hence one faraday is responsible for the movement of 1 mole of electrons. It is found to be equal to 96500 coulombs.

1 Faraday = 96500 coulomb

1 Faraday = 1 mole of es.

So 1 mole electrons carry 96500 coulomb electricity

6.023×10^{23} electrons carry 96500 coulomb electricity

i.e. 1 electron carry $\dfrac{96500}{6.023 \times 10^{23}}$ coulomb

$= 1.6 \times 10^{-19}$ coulomb/electron.

The Process of Electrolysis : In order to have a clear understanding of the process of electrolysis let us study the following reactions:

(a) **Electrolysis of water** :

$$2\,H_2O(l) \xrightarrow{\text{Electrolysis}} 2H_2(g) + O_2(g)$$

$$H_2O \rightleftharpoons H^+ + OH^-$$

At cathode

$H^+ + e \rightarrow H$ **Reduction**

$H + H \rightarrow H_2(g)$

At anode $OH^- - e \rightarrow OH$ **Oxidation**

$4OH \rightarrow 2H_2O + O_2(g)$

(b) Electrolysis of molten lead bromide

$PbBr_2 \rightleftharpoons Pb^{++} + 2Br^-$

At anode

$Br^- - e \rightarrow Br$ **Oxidation**

$Br + Br \rightarrow Br_2$

At Cathode $Pb^{++} + 2e \rightarrow Pb$ **Reduction**

(c) Electrolysis of NaOH

$NaOH \rightleftharpoons Na^+ + OH^-$

At anode

$OH^- - e \rightarrow OH$ **Oxidation**

$4\,OH \rightarrow 2\,H_2O + O_2$

At Cathode

$Na^+ + e \rightarrow Na$ **Reduction**

(d) Electrolysis of aqueous solution of $CuCl_2$ using copper cathode and carbon anode.

$CuCl_2 \rightleftharpoons Cu^{++} + 2\,Cl^-$

At anode $2Cl^- - e \rightarrow Cl$ **Oxidation**

$Cl + Cl \rightarrow Cl_2$

At cathode $Cu^{++} + 2\,e \rightarrow Cu$ **Reduction** (*e*) Electrolysis of molten Sodium chloride

$NaCl ; \Rightarrow Na^+ + Cl^-$

At anode $Cl^- - e \rightarrow Cl$ **Oxidation**

$Cl + Cl \rightarrow Cl_2$

At Cathode

$Na^+ + e \rightarrow Na$ **Reduction**

In molten NaCl, Na^+ and Cl^- ions are free to move when electrodes are dipped into molten NaCl and a potential is applied *i.e.* electricity is passed. Na^+ ion being positively charged move towards negative *cathode*. Here Na^+ ions lose its positive charge by gaining an electron to become a sodium atom.

At Cathode $Na^+ + e \rightarrow Na$

(It is a reduction reaction since *gains of e^-* is known as *reduction*).

A number of sodium atoms formed in this way collect at cathode and form molten sodium.

At the same time negatively charged chloride ions (Cl^-) are attracted towards *positive anode*. Here each chloride ion (Cl^-) lose an e^- and become chlorine atom. Two such chlorine at anode $Cl^- - e \rightarrow Cl$.

(It is an oxidation reaction since loss of e^- is oxidation).

atoms combine immediately to produce molecules of chlorine. So the reaction at electrodes can be summed up as :

Electrodes	Reaction	Type of Reaction
Anode	$Cl^- - e \rightarrow Cl$ $Cl + Cl \rightarrow Cl_2$	Oxidation
Cathode	$Na^+ + e \rightarrow Na$	Reduction

Such a process where an electrolyte is decomposed by an electric current is called Electrolysis. The electrolysis of molten electrolytes liberates metals at the cathode and non-metals at the anode.

6.5 MOVEMENT OF IONS AND ELECTRONS DURING ELECTROLYSIS

1. A solid electrovalent compound such as NaCl consist of ions held together by electrostatic forces of attraction.
2. In the molten state or inaqueous solution the ions move freely.
3. When electrodes are dipped in the electrolyte ions move towards cathode and anions move towards anode.

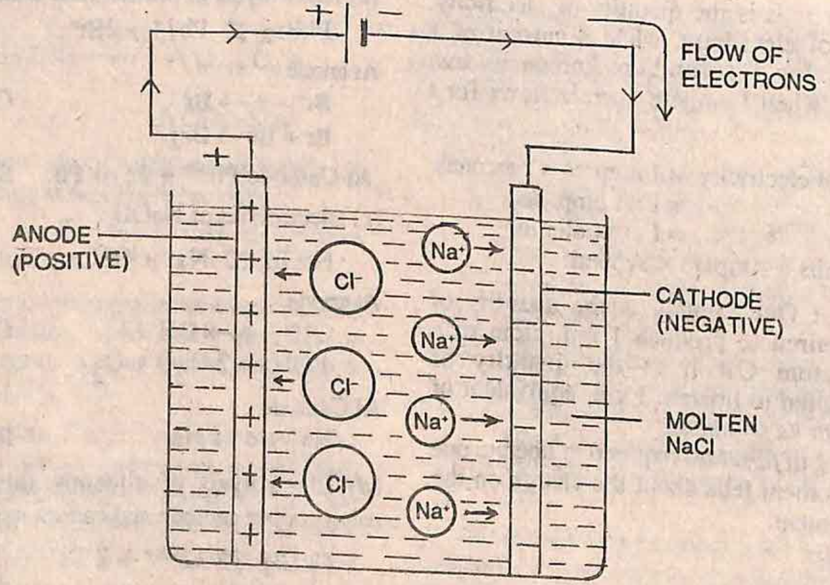


Fig. 6·6. Electrolysis of molten NaCl

4. This flow of ions is the electric current inside the electrolyte.
5. The ions transfer electrons to the anode and cations gain electron from the cathode.
6. The electrons lost by the anode equals the number of electrons picked up by the cathode. So the electrons move from anode through the battery to the cathode.
7. This flow of electron is the electric current outside the electrolyte.

6.6 QUANTITATIVE TREATMENT OF ELECTROLYSIS

In order to study the quantitative treatment of electrolysis let us take the example of the electrolysis of molten *NaCl*.

$$NaCl \rightleftharpoons Na^+ + Cl^-$$

At cathode $Na^+ + e \rightarrow Na^-$

At anode

$$Cl^- - e \rightarrow Cl$$

$$Cl + Cl \rightarrow Cl_2$$

So one electron changesone sodium ion (Na^+) to a sodium atom.

2 electrons change two sodium ion to two Na atom *n* electrons change n Na^+ to n Na atoms.

Hence N (Avogadro's No.) electrons change N Na^+ to N Na atom. The quantity of electricity carried by N electron *i.e.*, Avogadro *n* number of electrons is known as Faraday.

So one faraday of electricity when passed will give 1 mole Na atoms.

In the same way one faraday of electricity when passed will give 1 mole chlorine atoms or $\frac{1}{2}$ mole Cl_2 gas.

So when one faraday electricity is passed through molten NaCl ; 1 mole sodium is obtained and $\frac{1}{2}$ mole chlorine gas is obtained.

Let us take another example of electrolysis : *i.e.*, Electrolysis of molten $PbBr_2$

$$PbBr_2 \rightleftharpoons Pb^{++} + 2\,Br^-$$

$Pb^{++} + 2\,e \rightarrow Pb$ At cathode

$2\,Br^- - 2e \rightarrow Br_2$ At anode.

Hence to get 1 Pb atom 2 e^- are required.

To get 1 mole Pb atom 2 mole e^- are required or To get 1 mole Pb atom 2 faraday electricity is required.

To get 1 mole Br_2 2 faraday electricity is required.

Therefore to liberate *1 mole of any atoms* during electrolysis require an integral number of faradays depending upon the charge on the ion.

e.g.,	Na^+	require 1 faraday
	Mg^{++}	require 2 faraday
	Al^{+++}	require 3 faraday
	C^{4-}	require 4 faraday
	Pb^{+4}	require 4 faraday and *so on.*

Example 6.7 :

The charge on an electron is $1{\cdot}6021 \times 10^{-19}$ coulomb. If one faraday is equal to 96500 coulomb then calculate the total number of electrons that carry one faraday charge.

Solution:

1 faraday = 96500 coulomb

Charge on an electron = $1{\cdot}6021 \times 10^{-19}$ coulomb

So if the charge is 1.6021×10^{-19} coulomb then there is = 1 electron.

If the charge is 1 coulomb then there is

$$\frac{1}{1{\cdot}6021 \times 10^{-19}} \text{ electron}$$

If the charge is 96500 coulomb then there are

$$= \frac{96500}{1{\cdot}6021 \times 10^{-1_9}} \text{ electrons}$$

$$= \frac{96500 \times 10^{19} \times 10^4}{16021} \text{ electrons}$$

$$= 6{\cdot}023 \times 10^{23} \textbf{ electrons Ans.}$$

Example 6.8 :

Calculate the number of moles of copper deposited during the electrolysis of copper sulphate solution, when one faraday electricity is passed.

Solution:

During the electrolysis of copper sulphate, copper will be deposited at the cathode.

The cathodic reaction is $Cu^{2+} + 2\,e \rightarrow Cu$.

So for depositing 1 mole of copper the electricity required will be 2 faraday.

Hence when 2 faraday electricity is passed then 1 mole Cu is deposited.

So when 1 faraday electricity is passed then $\frac{1}{2}$ mole Cu is deposited = **0·5 mole copper. Ans.**

Example 6.9 :

During the electrolysis of molten calcium chloride following reactions take place at the electrodes $CaCl_2 \rightleftharpoons Ca^{++} + 2\,Cl^-$

At anode $2\,Cl^- - 2\,e \rightarrow Cl_2$

At cathode $Ca^{++} + 2e \rightarrow Ca$

(Atomic weight of Ca = 40) Calculate.

Amount of electricity in faradays and in coulombs that you will require to obtain :

(*a*) 1 mole calcium atom

(*b*) 4 gm of calcium atoms

(*c*) 1 mole of Cl_2 gas

(*d*) 11·2 litres of Cl_2 at S.T.P.

Solution :

(*a*) Calcium will be obtained at cathode.

The cathodic reaction is $Ca^{++} + 2\,e \rightarrow Ca$.

So 1 mole of Ca atom will require 2 mole electrons.

Electricity corresponding to 2 mole electrons is 2 faraday.

So 1 mole of Ca atom will require *2 faradays* of electricity. **Ans.**

1 faraday = 96500 coulombs

So to deposit 1 mole Ca we will require 2 × 96500 coulombs

= 193000 coulombs.

(*b*) 1 mole calcium atom = 40 gm calcium.

So for 40 gm. calcium we require 2 faraday electricity. For 4 gm. calcium we require $\frac{2 \times 4}{40}$ Faraday electricity

= 0·2 Faraday electricity

Ans.

For 40 gm calcium we require 1930000 coulomb electricity

4 gm calcium we require $\frac{1930000 \times 4}{40}$ coulomb electricity

= 193000 coulomb

Ans.

(*c*) Chlorine will be obtained at anode

The anodic reaction is $Cl^- - e \rightarrow Cl$

$Cl + Cl \rightarrow Cl_2$

So for getting 1 mole Cl_2 we will require 2 mole electrons.

OR for getting 1 mole Cl_2 we will require *2 faraday* electricity.

Ans.

For getting 1 mole Cl_2 we will require 2 × 965000 coulomb

= 1930000 coulomb

Ans.

1 mole of chlorine gas at S.T.P. occupy 22·4 lit. volume

OR 22·4 lit. of Cl_2 at S.T.P. = 1 mole Cl_2

Or 11·2 lit of Cl_2 at S.T.P. $\equiv \frac{11\cdot2}{22\cdot4}$ mole Cl_2

$= \frac{1}{2}$ mole of Cl_2 = *0·5 mole*

1 mole of Cl_2 will require 2 faraday electricity

0·5 mole of Cl_2 will require 2 × 0·5 = *1 faraday electricity*

Ans.

1 mole of Cl_2 will require 2 × 96500 coulomb electricity

0·5 mole of Cl_2 will require 2 × 96500 × 0·5

= *96500 coulomb*

Ans.

Electrochemical Cells: In the previous section we have studied the conversion of electrical energy to chemical energy and in the present unit we will study the conversion of chemical energy to electrical energy.

The electrodes, electrolyte solutions and container used in electrolysis experiments also constitute an electrochemical cell. The only question is, Is it possible to construct a cell in which chemicals react to produce an electric current instead of using current to decompose the electrolyte ? At the end of the eighteenth century Galvani and Volta.

6.7 ELECTROCHEMICAL CELL OR GALVANIC CELL OR VOLTAIC CELL

It is a device in which *chemical energy is converted to electrical energy*. It consists of two half cells. Each half cell is made up of a metal and its electrolyte. The electrolytes in the two half cells are connected either by a porous pot or a salt Bridge (an inverted U tube). At one half cell oxidation occurs and at the other half cell reduction occur. The electrode where *oxidation* occur is known as *anode* and the electrode where *reduction* occur is known as *cathode*. In case the two electrodes are Zinc and Copper then it is known as Daniel Cell.

Competition for Electrons : The half cell reactions in the Daniel cells are :

$Zn\,(s) - 2e \rightarrow Zn^{++}\,(aq)$ **Anode**

$Cu^{++}(aq) + 2\,e \rightarrow Cu(s)$ **Cathode**

These half reactions can be viewed as a competition between two kinds of metal atom for electrons. In this case zinc atoms being more reactive have greater tendency to lose electrons as compared to copper.

$$Zn(s) + Cu^{++}(aq) \rightarrow Zn^{++}(aq) + Cu(s)$$

In this way the metals can be arranged in the order of their reactivities as follows :

Sodium > Magnesium > Aluminium > Zinc > Iron > Lead > Hydrogen > Copper> Mercury > Silver > Gold. This arrangement of the elements in the decreasing order of their reactivities constitute. electrochemcial series.

Electrochemical Series : In the electrochemical series the metal that comes first is more reactive and has a greater tendency to form positive ions. The more active metal can displace the less active metal from the solution. So with the help of electrochemical series we can :

(*i*) *Predict feasibility of chemical reactions:* Suppose we want to know whether any reaction will occur if an iron knife is dipped in a solution of copper sulphate. If iron lie above copper in the series then iron is more reactive than copper and has a tendency to form positive ions and the reaction takes place. It is due to this reason copper sulphate cannot be stored in an iron container.

(*ii*) *Choose anode and cathode:* Out of the two metals that metal is taken as anode which has greater tendency to form positive ion, *e.g.*, out of Fe and Cu, Fe is anode and copper is cathode because iron comes before copper in the electrochemical series.

(*iii*) Predict why most of the metals are found in the form of ionic compounds in nature. This is because metals that are before hydrogen generally go to ionic form and exist in the form of compounds.

6.8 CELLS IN EVERYDAY USE :

The cells that we have studied so far have two disadvantages :

(1) They are not easily recharged.

(2) They are difficult to transport because they contain liquids. The two most important cells that are used made it possible. Volta produced electricity by the chemical action between two different metals dipped in their electrolyte and he made the first practical battery. The two metals used by him were Zinc and Copper.

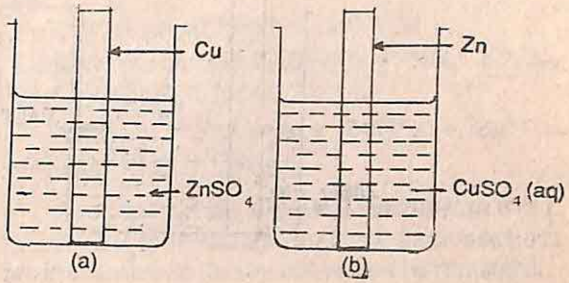


Fig. 6·7

Take two beakers A and B. In the beaker take Zn SO_4 solution and dip a copper rod in it and in the beaker B take copper sulphate solution and dip zinc rod in it. Nothing happens in the beaker A whereas in beaker B, colour of copper sulphate solution fades away and a reddish brown sludge of copper is deposited.

Equations :

In beaker A $Cu(s) + Zn^{2+} SO_4{}^{2-} (aq) \rightarrow$ No Reaction.

In beaker B $Zn(s) + Cu^{2+} SO_4{}^{2-} (aq) \rightarrow Zn^{2+} SO^{2-}{}_4(aq) + Cu(s)$.

In the beaker B each zinc atom loses 2 electrons to each copper ion in the solution

$$Zn^{-2e^-} \rightarrow Zn^{2+}$$

$$Cu^{2+} + 2e \rightarrow Cu$$

Each of these reactions are known as *Half Reactions.*

The half reaction in which electrons are lost is known as *oxidation* and the half reaction in which electrons are gained is known as *Reduction.*

Although there is a transference of electrons but no electric current is produced because Zn and Cu^{++} are in contact with each other and electrons do not move through a considerable distance. Hence in this case chemical energy is converted to heat energy. The electric current is produced only when the electrons are made to flow through a wire. A device for creating a flow of electrons through the wire is known as *primary cell.*

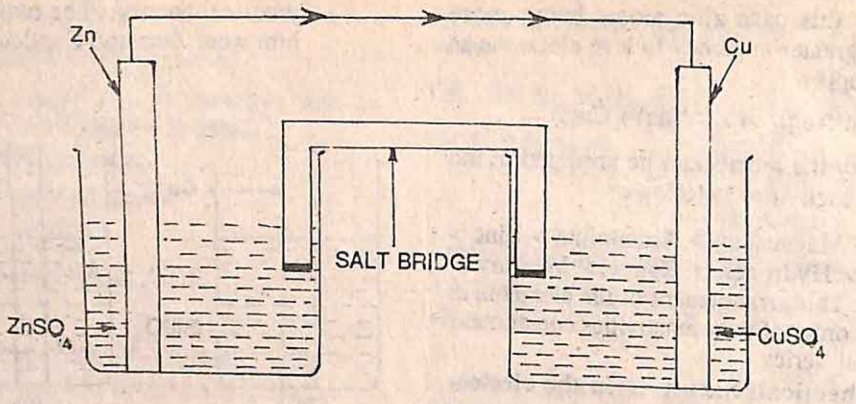


Fig. 6·8. A Primary Cell

Conventions of Electrode	Cathode	Anode
Movement of Electrons	Cations	Anions
Direction of electron flow	Into the cell	Out of the cell
Half Reaction	Reduction	Oxidation
Sign of electrode in		
(*a*) Electrochemical cell	Positive	Negative
(*b*) Electrolytic cell	Negative	Positive

So an arrangement is made for the indirect flow of electrons. It consists of two beakers filled with the electrolytes. The two electrolytes must be in contact with each other but are not allowed to mix. An arrangement as shown below is made :

In this case electrons move from zinc rod through the wire to copper rod and then to Cu^{2+} ions. The reactions taking place in the two beakers are:

$Zn \rightarrow Zn^{++} + 2e$ **Oxidation**

$Cu^{++} + 2e \rightarrow Cu$ **Reduction**

In this redox reactions electrons move through a considerable distance and flow of electron is flow of current. Hence a Primary Cell is made up of :

(*i*) A metal dipped in its own electrolyte *e.g.* , Zn in $ZnSO_4$ solution.

(*ii*) Another metal dipped in its own electrolyte *e.g.* , Cu in $CuSO_4$ solution.

(*iii*) The two metals are connected by a wire.

(*iv*) The two electrolytes are connected by an inverted U-tube filled with saturated salt solution (is called *Salt Bridge*) OR the electrolytes are connected with the help of a **Porous Pot.** Such a cell is called **Daniel Cell.** Each metal dipped in its electrolyte constitute a *half cell.*

Hence in *Daniel Cell* the Zinc rod gradually dissolves as atoms of Zinc lose electrons and become Zn^{++} ion. These electrons pass through the external circuit to copper rod where copper ions from copper sulphate solution are converted to copper metal. A voltmeter if connected across the terminals of the cell will register 1.1 volt. *This voltage recorded* when no current flows is known as *Electromotive force* of the cell.

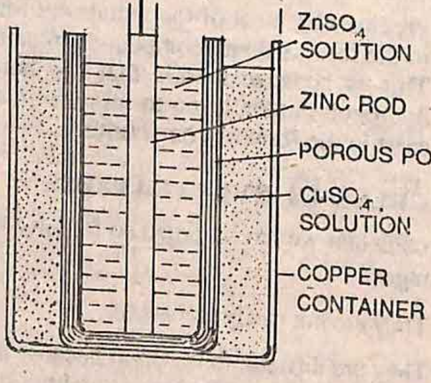


Fig. 6·9. Daniel Cell

By using different combinations of metals in a suitable electrolyte *i.e.*, solution of the metal salt, we can construct a variety of electrochemical cells giving different voltages in daily life also suffer from one of these difficulties.

Dry Cell Or Torch Battery : There are various types of dry cells but one of most well known is the common flash light battery or Leclanche dry cell. Dry cells generally contain an electrolyte in a paste rather than a liquid solution of the electrolyte – hence the name is *dry cell*. These are known as primary cells because it can be used once only and cannot be recharged.

Construction : In the centre of the cell is a graphite rod that serves as the positive electrode. This is surrounded by a paste of manganese dioxide (MnO_2) and graphite. [The purpose of manganese dioxide is to oxidise hydrogen produced at the electrode to water. Thereby it prevents the hydrogen gas from adhering to graphite which would otherwise reduce its efficiency]. The paste of manganese dioxide and graphite is further surrounded by a paste of Ammonium chloride. The rod and the paste are contained in a zinc case which acts as the negative electrode.

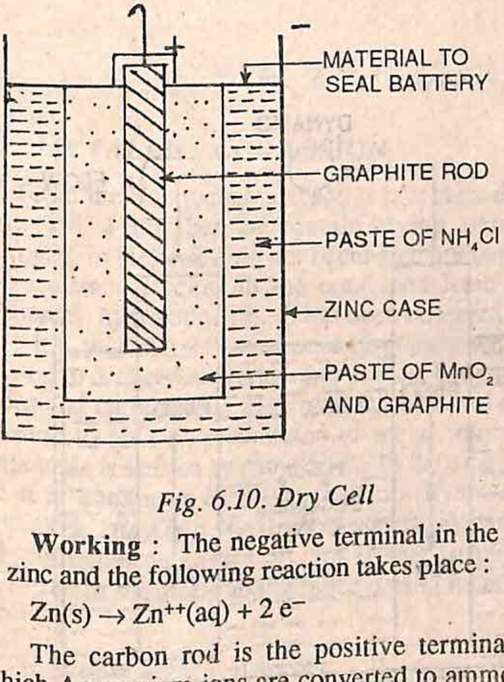


Fig. 6.10. Dry Cell

Working : The negative terminal in the cell is zinc and the following reaction takes place :

$$Zn(s) \rightarrow Zn^{++}(aq) + 2\,e^-$$

The carbon rod is the positive terminal at which Ammonium ions are converted to ammonia and hydrogen

$$2\,NH_4^+(aq) + 2\,\bar{e} \rightarrow 2\,NH_3(g) + H_2(g)$$

Hydrogen is oxidised by ammonia to H_2O.

$$2\,MnO_2(s) + H_2(g) \rightarrow Mn_2O_3(g) + H_2O(l)$$

The ammonia gas causes no problem because it is highly soluble in water of the paste. It is necessary to use paste because dry $NH_4\,Cl$ would not conduct electricity. *Hence* over all reactions are:

Anode [Oxidation] $Zn(s) \rightarrow Zn^{++} + 2\bar{e}$

Cathode [Reduction] $2\,MnO_2(s) + 2NH_4^+ + 2e^- \rightarrow Mn_2O_3(s) + H_2O + 2NH_3$

$$Zn(s) + 2\,MnO_2(s) + 2NH^+_4 \rightarrow Zn^{++} + Mn_2O_3(s) + H_2O + 2NH_3$$

As a result of the chemical reactions an e.m.f. of 1·5 volts is produced. Since all these reactions are irreversible so the dry cell cannot be recharged.

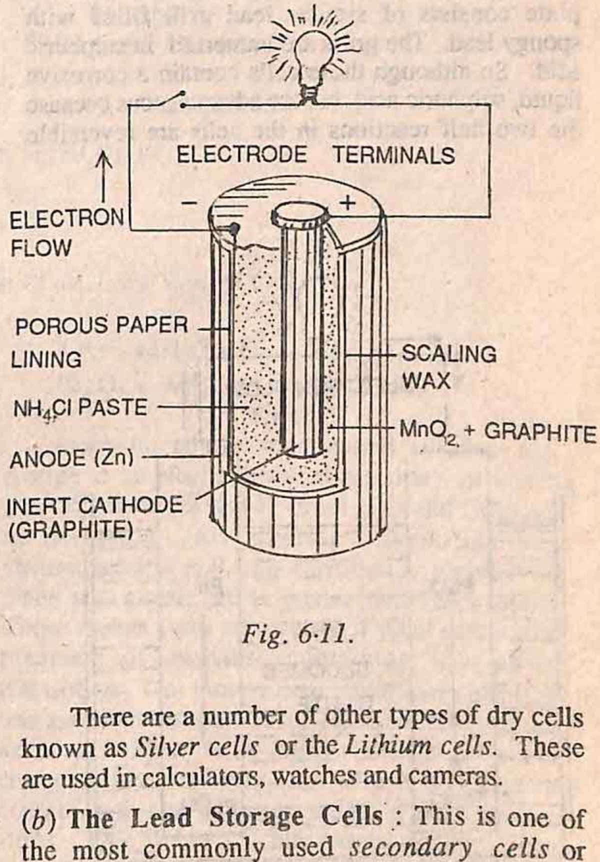


Fig. 6·11.

There are a number of other types of dry cells known as *Silver cells* or the *Lithium cells*. These are used in calculators, watches and cameras.

(*b*) **The Lead Storage Cells** : This is one of the most commonly used *secondary cells* or *accumulators*. A secondary cell is one that can be recharged and used again. All car batteries are composed of either three or six lead acid cells in series and give a voltage of nearly 6 or 12 volts respectively.

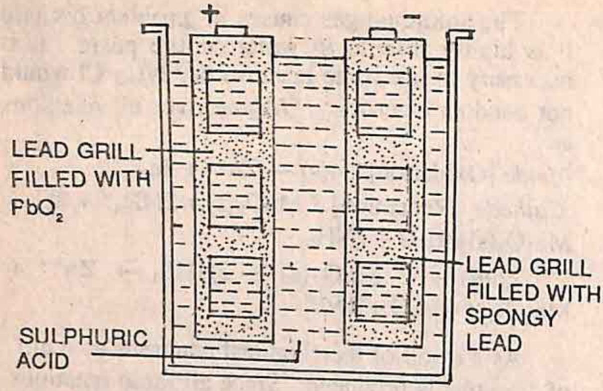

+
−
LEAD GRILL FILLED WITH PbO_2
LEAD GRILL FILLED WITH SPONGY LEAD
SULPHURIC ACID

Fig. 6.12. A part of lead storage battery.

Construction : The positive plate is lead grill filled with lead dioxide (PbO_2). The negative plate consists of similar lead grill filled with spongy lead. The grills are immersed in sulphuric acid. So although these cells contain a corrosive liquid, sulphuric acid, but are advantageous because the two half reactions in the cells are reversible hence can be recharged.

Working : The negative terminal is lead which gives up electrons forming lead ions (Pb^{++})

$Pb(s) \rightarrow Pb^{++} + 2\,\bar{e}$ *Anode Oxidation*

The positive terminal is lead dioxide that reacts with hydrogen ions, H^+, in sulphuric acid electrolyte.

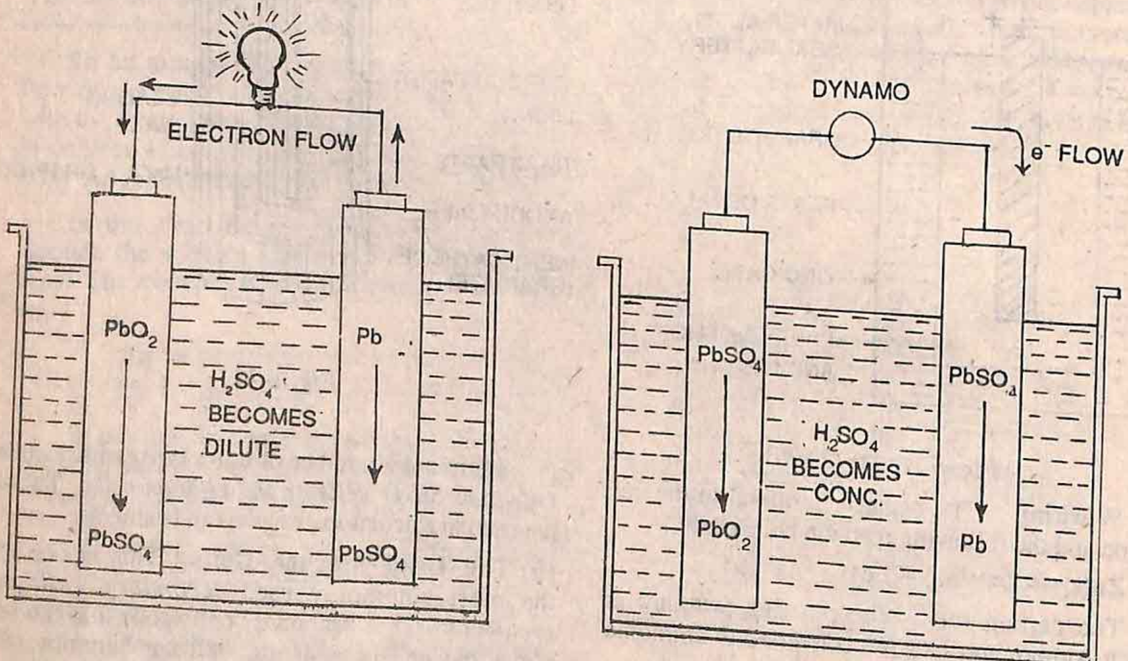
$PbO_2(s) + 4\,H^+(aq) + 2\bar{e} \rightarrow Pb^{++}(aq) + 2\,H_2O(l)$ *Cathode Reduction.*

The lead ions formed during the reaction reacts with sulphate ions and form insoluble lead sulphate

$Pb^{++}(aq) + SO_4^{--}(aq) \rightarrow PbSO_4(s)$

The specific gravity of the sulphuric acid is a measure of the effectiveness of the recharge. Because the density of the sulphuric acid is greater than that of water. The cell reaction is reversible so by applying a potential to the battery in excess of its discharge potential we can charge the battery. Hence during charging we want to carry out an electrolysis reaction that converts the electrodes back to the original materials. The net reaction is

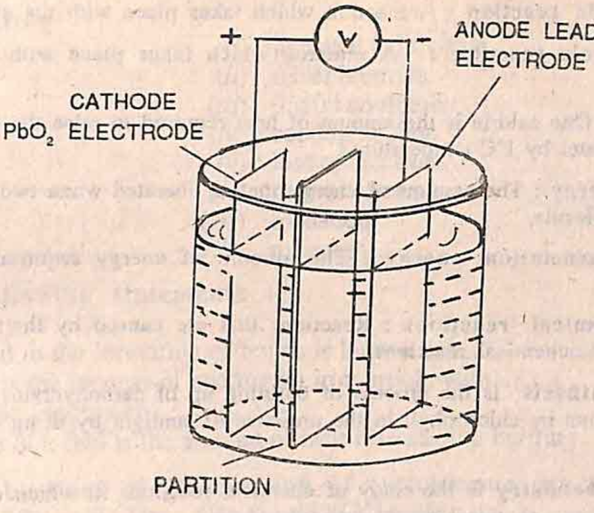
$2\,PbSO_4(s) + 2\,H_2O(l) \rightarrow Pb(s) + PbO + 4H^+(aq) + SO_4^{--}\,(aq)$



Reactions: During Discharge

During Charge

Fig. 6.13 Charging & discharging processes in the Lead storage battery.



ELECTRODE REACTION
(ANODE)
$Pb(S) \longrightarrow Pb^{++} + 2e^-$

ELECTRODE REACTION
(CATHODE)
$PbO_2(S) + 4H^+ + 2e \longrightarrow Pb^{++} + 2H_2O(l)$

Fig. 6.14 Schematic Diagram of the Lead Storage Cell.

6.9 METALLIC CORROSION

From the electrochemical series it is clear that there are a number of metals, lying above hydrogen, in the series, do not occur free in nature. These elements occur in the combined form as carbonates, hydroxides etc. It has been observed in daily life that metals like iron and copper when exposed to air gets converted into oxides, hydrated oxides and carbonates. This wastage of metal on exposure to air *i.e.*, conversion of metal into its compounds is known as *Corrosion*. In the case of iron it is known as *Rusting*. Rust is hydrated ferric oxide ($Fe_2O_3 . x H_2O$). It is non-sticky and peals off exposing fresh iron for further rusting. The different reactions taking place during rusting are :

$Fe \rightleftharpoons Fe^{++} + 2\bar{e}$ [Anode]
$H^+ + e \rightarrow H$ [Cathode]
$4H + O_2 \rightarrow 2H_2O$
$4Fe^{++} + 4H_2O + O_2 \rightarrow 2Fe_2O_3 + 8H^+$
$Fe_2O_3 + xH_2O \rightarrow Fe_2O_3 . X H_2O$
Rust

From the above reactions it is clear that rusting is an electrochemical reaction. In which *iron* acts as *anode* and H^+(from H_2O and CO_2) act as *cathode*. So rusting constitute an electrochemical cell. So corrosion is speeded up when two metals are in contact with each other. These metals form an electrochemical cell in the presence of moisture containing CO_2 as an electrolyte. The more active metal loses electron and goes into the ionic state. Hence if iron rods are kept in contact with copper or tin then the corrosion becomes very fast. But if it is kept in contact with zinc or magnesium then the rusting stops. This is used in preventing rusting. Iron surface is covered with a layer of zinc. The process of covering iron with zinc is known as galvanisation. Galvanised iron sheet maintain the lustre due to the formation of Basic Zinc carbonate [Zn CO_3. $Zn(OH)_2$]. This type of coated iron is called galvanised iron or G.I.

SUMMARY

6.1 **Exothermic reaction** : A reaction which takes place with the evolution of heat.

6.2 **Endothermic reaction** : A reaction which takes place with the absorption of heat.

6.3 **Calorie** : One calorie is the amount of heat required to raise the temperature of one gram of water by 1°C temperature.

6.4 **Bond Energy** : The amount of energy that is liberated when two atoms combine to form a molecule.

6.5 **Bond dissociation energy** : The amount of energy required to break up the bond.

6.6 **Photochemical reactions** : Reactions that are caused by the action of light are called photochemical reactions.

6.7 **Photosynthesis** is the process of building up of carbohydrates in the green parts of the plants by chlorophyll in the presence of sunlight by using carbon dioxide and water.

6.8 **Electrochemistry** is the study of chemical reactions in which electrical energy is converted into chemical energy and vice versa.

6.9 **Electrolytic cells** are those in which electrical energy is converted to chemical energy.

6.10 **Electrochemical** cells are those in which electricity is produced at the cost of chemical reactions.

6.11 **Electrolyte** is a compound which when melted or in solution conducts an electric current and is decomposed by it.

6.12 **Non-electrolyte** is a compound which when melted or in solution does not conduct an electric current.

6.13 **Electrolysis** is the decomposition of an electrolyte by passing an electric current.

6.14 **Electrodes** are the two pieces of metal or graphite by which electric current enters and leaves an electrolyte.

6.15 **Cathode** is the negative electrode by which electrons enter an electrode.

6.16 **Anode** is the positive electrode by which electrons leave the electrode.

6.17 **Ampere** is the unit for measuring the flow of electrons in a circuit.

6.18 **Coulomb** is the quantity of electricity when 1 ampere current flows for one second.

6.19 **Faraday** is the quantity of electricity required to produce 1 gm atom of a monovalent atom.

6.20. **Oxidation** is the loss of one or more electrons by an atom or a group of atoms or ions.

6.21. **Reduction** is the loss of one or more electrons by an atom, a group of atoms or ions.

6.22. **Electrochemical** series is the arrangement of the elements in the decreasing order of their reactivity.

6.23. **Corrosion** is the wastage of metal on exposure to air.

6.24. **Rusting** is the wastage of iron on exposure to air.

6.25. **Rust** is hydrated ferric oxide (Fe_2 O_3. xH_2O).

6.26. **Galvanisation** is the process of covering iron with zinc.

QUESTIONS

6.1 Match the following :

(*i*)	Faraday	(*i*)	hydrated ferric oxide
(*ii*)	Galvanisation	(*ii*)	silver bromide
(*iii*)	Oxidation	(*iii*)	96500 coulomb
(*iv*)	Reduction	(*iv*)	coating of Zinc
(*v*)	Anode	(*v*)	loss of electrons
(*vi*)	Cathode	(*vi*)	gain of electrons
(*vii*)	Photography	(*vii*)	oxidation
(*viii*)	Rust	(*viii*)	reduction

6.2 Complete the following statements :

(*a*) Units of energy is ....

(*b*) Energy evolved in the formation of bonds is known as.............

(*c*) Galvanisation is the process of coating an iron article with ....

(*d*) Oxidation is the process of ....

(*e*) Calorific value of a fuel is the amount of heat liberated by burning .... gms. of a fuel.

6.3 The heat of combustions of glucose and of cyclohexane are given below :

(*a*) $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O + 2825$ KJ/mole

(*b*) $C_6H_{12} + 9O_2 \rightarrow 6CO_2 + 6H_2O + 3950$ KJ/mole

Which would give more energy per gram ?

6.4 "The heat of combustion of butane is 2890 KJ Mole^{-1}". Explain the meaning of this statement. And calculate the amount of heat liberated by the complete combustion of 1 gm of butane.

[Butane C_4H_{10} : C = 12; H = 1]

6.5 Define the terms : Bond energy and Bond dissociation energy. Which of the two is an exothermic process ?

6.6 On heating mercuric oxide decomposes to give mercury and oxygen, according to the equation

$$2HgO + 180\ KJ \rightarrow 2Hg + O_2.$$

Calculate the heat of formation of mercuric oxide. Also calculate the amount of heat needed to decompose one mole of mercuric oxide.

6.7 Which of the following reactions would you expect to be exothermic and which to be endothermic ?

(*a*) $Cl_2(g) \rightarrow 2Cl(g)$

(*b*) $Na(s) \rightarrow Na(g)$

(*c*) $C(s) + O_2(g) \rightarrow CO_2(g)$

(*d*) $Na(g) \rightarrow Na^+(g) + e^-$

(*e*) $Na^+(g) + Cl^-(g) \rightarrow NaCl(s)$

6.8 Calculate the amount of heat liberated during the formation of hydrogen chloride from the following thermochemical equations :

(*a*) $H_2(g) + 433\ KJ \rightarrow 2H(g)$

(*b*) $Cl_2(g) + 239\ KJ \rightarrow 2Cl(g)$

(*c*) $H(g) + Cl(g) \rightarrow HCl(g) + 433\ KJ$

6.9 One faraday carry charge equal to one mole of electrons. Calculate the charge in coulombs present on one electron.

6.10 Write down the reactions taking place at the anode and cathode during the electrolysis of

(*a*) sodium hydroxide

(*b*) molten sodium chloride

(*c*) molten lead bromide.
Also name the elements oxidised and reduced.

6.11 Calculate the number of moles of lead deposited during the electrolysis of molten lead bromide, when one faraday electricity is passed.

6.12 Below is given the information regarding three metals X, Y and Z.
(*a*) Z displaces X from its aqueous solution.
(*b*) Z is obtained by the reduction of its oxide with coke.
(*c*) Y cannot be obtained by the reduction of its oxide with coke but can be obtained by the electrolytic reduction of its molten salt.
Arrange X, Y and Z in the decreasing order of its reactivity.

6.13 A, B and C are the three elements that undergo chemical change according to the following equations :
(*i*) $A_2O_3 + 2B \rightarrow B_2O_3 + 2A$
(*ii*) $3\,CSO_4 + 2\,B \rightarrow B_2(SO_4)_3 + 3C$
(*iii*) $3\,CO + 2A \rightarrow A_2O_3 + 3C$
Arrange these elements in the increasing order of reactivity.

6.14 Why colour of copper sulphate solution is discharged after sometime, when iron rod is dipped into it ?

6.15 Define rusting. State one similarity between
(*a*) rusting and burning.
(*b*) rusting and respiration.

6.16 Why is iron galvanised with zinc
Can it be galvanised with copper or lead ? If not why not ?

6.17 What is rust ? Why is an iron grill painted ?

6.18 List the conditions necessary for rusting. How can you protect iron from rusting ?

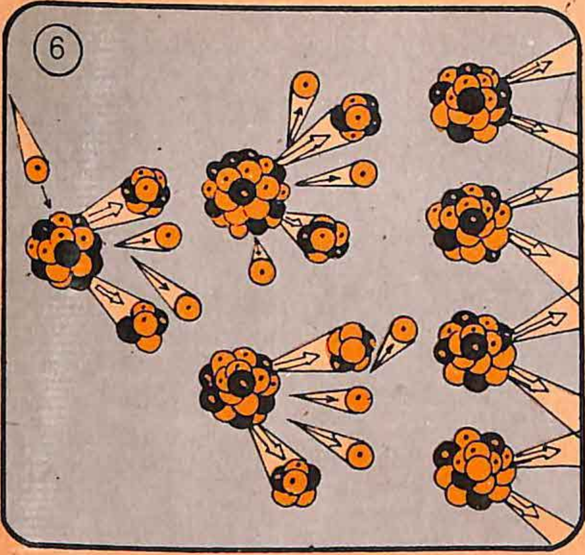
6.19 Draw a well labelled diagram of a Daniel Cell.

6.20 What are primary cells ? Why are they so called ? Is dry cell a primary cell or a secondary cell ?

6.21 Name the type of cells used in cars. Why are they called secondary cells ? Name the anode, cathode and electrolyte in these cells.

6.22 Write down the reactions taking place during the charging and discharging of lead storage cells.

□□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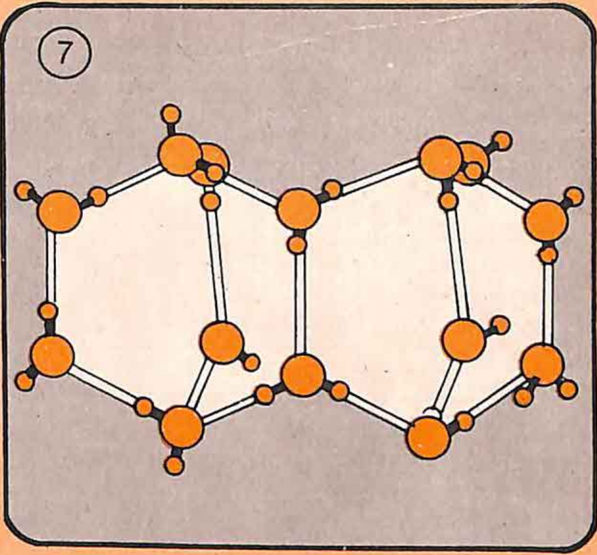
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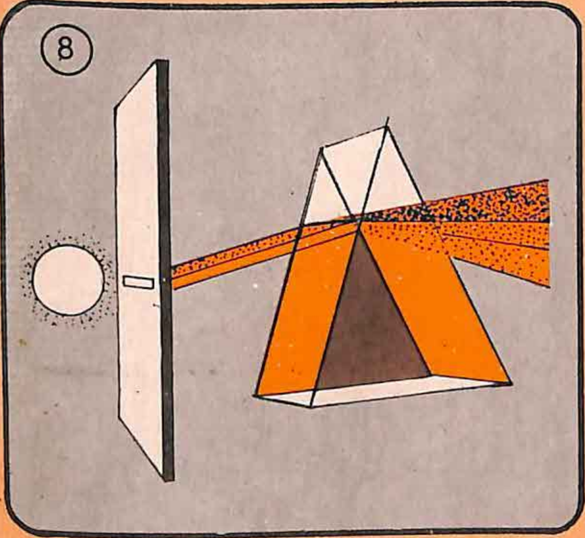
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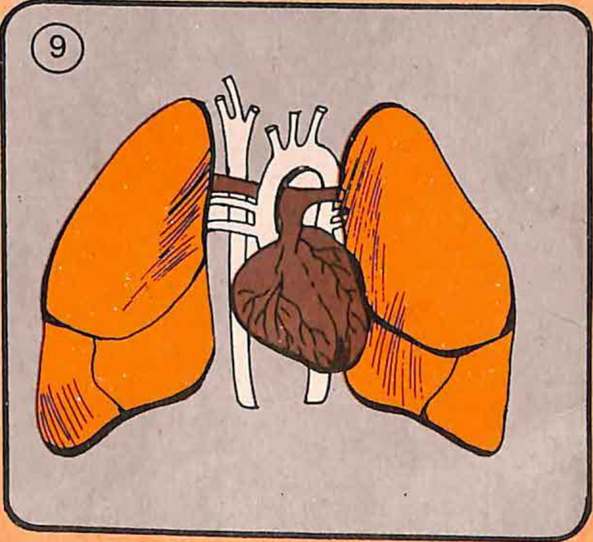
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